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THE

PHASIS OF MATTER.

VOL. I

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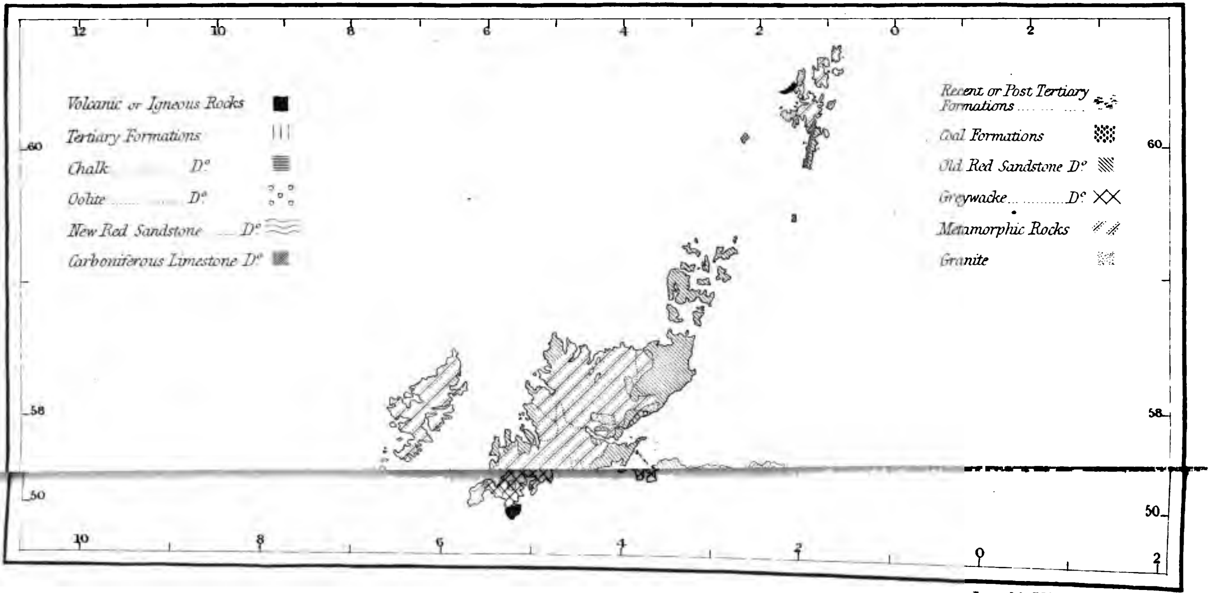

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THE
PHASIS OF MATTER;

BEING AN

OUTLINE OF THE DISCOVERIES AND APPLICATIONS
OF
MODERN CHEMISTRY.

BY T. LINDLEY KEMP, M.D.

> "Omnia mutantur; nihil interit. Errat et illinc
> Huc venit, hinc illinc." Ovid.

IN TWO VOLUMES.

VOL. I.

LONDON:
LONGMAN, BROWN, GREEN, AND LONGMANS.
1855.

PREFACE.

Two objections, I fear, may be made to the publication of these volumes. It may be said that there is no need of another treatise upon chemistry; and, in the second place, that if there is, it should be written by a professed chemist, and not by one who, like myself, has only attended to chemistry in common with physiology and other sciences.

To the first of these objections I would say, that there certainly are in our language more than two or three very excellent manuals of chemistry, but all written so as to suit the wants of those who intend to make chemistry the main object of their lives. There are also one or two good elementary books for the use of schools. But it appears to me that we do not possess a treatise on chemistry more extended than the latter, but less minute than the former, and intended for the wants of the general scholar and of men of the world, whose active occupations are more or less based upon a knowledge of chemical principles

Keeless. MLP, 5-10-37

and chemical facts. And yet chemistry has now become of such importance, that every educated man, to be such, must have a general knowledge of it; and there are a great number of professions and trades, of which medicine, farming, calico printing, or brewing may be instanced at random, which can only be satisfactorily conducted, and be reasonably expected to be improved, by means of chemistry. To supply this numerous class with a manual of chemistry, having a moderate bulk and price, has been my object in writing the following sheets.

With regard to the second of these objections, I am not sure if, considering my intentions respecting the book, that position of mine which I suppose to be complained of be not a positive advantage. There is, in truth, but *one* science, and our artificial divisions of it have sometimes surely had their barriers too well kept up. Man is always inclined to be pedantic, and the isolated ideas of the scholastic pedagogue have, perhaps, insensibly crept into the domain of science. At least, the reason that so very many students of chemistry, certainly, the majority, break down in despair, long ere they have mastered the subject, seems to me to be partly owing to its being treated so much as an abstract matter by itself, and so little in common with the general philosophy of matter.

I have tried to go upon another plan, and have endeavoured to dwell upon the fact, that the student of chemistry must not be content with the mere chemistry of the laboratory. This, indeed, is absolutely necessary to know, and is the basis of all farther inquiry; but a great part of it is, in one sense, quite artificial.

Accordingly, in the first Book I attempt to state the Properties of the Elementary Bodies, and the compounds that they form when two or more are put together in a pretty isolated state in the laboratory. But it is notorious, that in the laboratory of nature these elements form much more complex compounds. Hence, I have thought it necessary to notice in Book II. the chemical nature of the Compounds that naturally exist in the Minerals and Rocks of the Globe. In Book III. I endeavour to describe the Chemistry of Animal and Vegetable Bodies that are only just dead; or, to use the common expression, Organic Chemistry. In the next division I attempt an outline of what seems to be most probable with regard to the Changes that Matter undergoes in these bodies before they die, and when their vitality is in full vigour; and I conclude with an account of some of the Applications of Chemistry to the Arts.

The subject is a very wide and a difficult one;

and I cannot hope that the following pages do not contain many faults, both of omission and commission. Still, I have bestowed labour upon it, and honestly tried to make it a System of Chemistry that may be of use to the general scholar and man of the world. And if I have at all succeeded in this, I shall not grudge the pains.

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THE
PHASIS OF MATTER.

INTRODUCTION.

We can easily conceive that, at an early period of human society, no one had discovered any of the truths or principles of science. Indeed, at the present day, and in the most highly civilised countries of Western Europe, the great majority of mankind are quite ignorant of such; and there are yet whole nations who do not know one of the revelations or abstractions of science. Mankind, however, cannot exist without the exercise of some of the arts; and all tradition and all history tell that the two earliest of these that were cultivated were, first hunting, and subsequently farming. There is every reason to believe that for long man was content with the exercise of these two arts, with a little mechanical skill in making weapons and implements of wood and stone, and with manipulating skins of animals for raiment. The next art that was discovered, probably after a long interval, was that of metallurgy.

Even for a long time after the discovery of iron, and of the art of obtaining it from the ore, there does not appear to have been much, if any, division of labour. The father of the family, in those patriarchal times, was indeed the prince and the priest; but he seems to have toiled at the various necessary labours along with his children and dependants. Men engaged in daily physical drudgery cannot be expected to find out the principles of science, and the men to whom we now refer, doubtless, discovered none of them.

At length a sacerdotal class arose, the office of the members of which was to administer the rites of religion, and who were exempt, and indeed prohibited, from bodily toil. It was from the contemplations and reflections of this sacerdotal class that science took its origin. This modification of the original state of society appeared in Egypt, probably nearly four thousand years ago; and there can be no doubt but that the Egyptian priests arrived at a great familiarity with the truths of science. Unfortunately, they took as much trouble to keep their discoveries to themselves as to make them, and Egyptian learning nearly perished with the extinction of the Egyptian monarchy and religion. We know, however, that they were well acquainted with astronomy; their pyramids, by far the vastest monuments of stone that the world has ever possessed, attest an intimate familiarity with geometry and mechanics; the word *chemeia*, from which chemistry is derived, is an Egyptian expression; and it is almost certain, that as long ago as six centuries before the Christian era, the Egyptians were aware of the existence of the Americas.

But although their knowledge mainly perished with them, it is likely that we have been indirectly affected by it, through the medium of Pythagoras and of Moses. The latter, bred up amongst the sacerdotal class, was "skilled in all the learning of the Egyptians;" and among the early Hebrew writings may be found traces of a great scientific principle, unknown to our ancestors and only just acquired by us. With regard to Pythagoras, so much mystery surrounds both his life and his doctrines, that it is not always easy to speak decidedly regarding either. He obtained admission into the college of Egyptian priests at Thebes, where he remained for many years. At length, when he had fully studied the Egyptian philosophy, he established himself as the head of a philosophical sect at Crotona. Much of what he taught was, perhaps, his own deduction; but so much at least must have been acquired from his instructors. Unfortunately he adopted an exoteric and an esoteric mode of teaching. To the many he addressed discourses, recommending temperance and control over the passions, applauded virtue, and rebuked vice. But he had a chosen band, only admitted after a long probation, to whom, and to whom alone, he delivered his system of philosophy. It is probable, indeed, that the open avowal of any philosophical doctrines was, in the days of mythological religion, dangerous; farther, Pythagoras committed nothing to writing, and we have to trust to subsequent authors for a knowledge of his doctrines. This renders an account of his opinions still more difficult.

One of the aims of Pythagoras would appear to have been to form a general system of philosophy, the various portions of which, the arithmetic, the geometry, the astronomy, the physics, the fine arts, the morals, and the theology, should have a mutual relation and dependence on one another, and should possess intimate analogies. His leading idea would seem to have been that every thing is in a state of eternal change and alteration from the state that it is in, but that eventually it will return to the same or to an analogous state to that which it starts from; and farther, that in many cases these changes take place in harmonic ratios that can be expressed by numbers.

Thus he taught that the sun was the centre of our system, and that the planets were for ever revolving around that luminary. He held that each planet was perpetually altering its position with regard to the sun, and yet always tending to regain its former position. For two thousand years this account of the planetary motions was regarded as false and chimerical. Since the time of Copernicus it has been, as it now is, universally believed to be the true statement.

He farther taught, that the relative distances of the planets from the sun were such, that the numbers expressing them would have the same relation to each other as the vibrations of a musical string in a harmonic chord. Into this subject we need not here enter.

The discovery of the principles of music is, on sufficient authority, attributed to Pythagoras. He found experimentally that, when successive notes followed one another, and when the numbers of the vibrations produced in the string during a second of time, were in harmonic ratios, as for example, when a tonic dominant and mediant accompany each other, there was harmony; and when this was not the case, that there was discord. It is almost certain that he held, that analogous measurements existed with regard to beauty in form. The accuracy of his dogmata regarding music is now universally admitted; but his hypothesis respecting other kinds of material beauty, have been for centuries disregarded. Of late they have been advocated, or rather again promulgated by Mr. Hay, and there is every prospect that they will eventually be received.

It is also curious to remark, that the doctrine now universally maintained by geologists, to the effect that the great masses of matter, lands and seas, are in a constant, although slow, state of transmutation into one another, was one of the doctrnes of Pythagoras. It is not, however, until within these last twenty years that this opinion has been accepted among the moderns.

It is difficult to decide as to the Pythagorean doctrine of the nature of matter; but it is probable that he held, that each mass of matter was made up of a congeries of atoms, or monads, as he termed them. We have no evidence, however, that Pythagoras applied his doctrine of numbers to the different combinations of various kinds of atoms in any way resembling the modern doctrine of chemical equivalents, or the atomic theory.

But perhaps the most remarkable of the opinions of Pythagoras was, that regarding the circulations of the different atoms of matter; and he held that they were continually changing their position. In the words of his commentator: —

> Omnia mutantur; nihil interit. Errat et illinc
> Huc venit, hinc illinc.

We, from our exacter knowledge of the chemistry of the elements, can express this doctrine more definitely. We see, for example, that a wheat plant is composed of the different elements called carbon, oxygen, iron, phosphorus, sodium, &c., and that the air and the soil around it contain these very elements, and afford them to the growing plant. In fact, we plainly perceive these elements passing from the inorganic world into the structure of the wheat plant. We then witness an animal whose structure is likewise composed of carbon, oxygen, iron, phosphorus, sodium, &c., eat this wheat plant, and the elements of the plant become the elements of the animal. We then farther witness the animal's death, the consequent putrefaction of its body, and the return of these elements to the soil, and the air to be again the food of a wheat or some other plant. Matter is for ever passing from the inorganic world to the organic, and back again from the organic to the inorganic. This doctrine can only be clearly expressed by one familiar with the results of modern chemical analysis. But it is plain that Pythagoras had at least an outlined idea of the phasis of matter.

Other causes besides his secret mode of teaching, contributed to prevent the philosophy of Pythagoras from obtaining a firm hold upon his contemporaries and their descendants. He and his college were mixed up with the politics of the day; and having joined the oligarchical faction, the populace rose upon them, burnt their residence, put many of them to death, drove the remainder into exile, and their great leader died of hunger in his flight. Moreover, at this period knowledge had arrived at such a stage in her progress that she began to be divided into sections, and one man devoted himself to mathematics, another to medicine, a third to astronomy, and so forth. One result of this was, that the desire of having a general system of philosophy that would embrace universal knowledge, became weakened. Another result would doubtless be a great addition of new facts; and in tracing the history of science, we can perhaps also notice that there are certain stages in the accumulation of facts that tend to produce the spirit of scepticism and that of encyclopedism.

Accordingly, in less than a century after the death of Pythagoras, we have the type of the former in Socrates, and a generation later, of the latter in Aristotle. The favourite and constantly reiterated saying of Socrates was, "He knew nothing save this — that he knew nothing," which is tantamount to a declaration of his distrust of the powers of the human mind to arrive at a knowledge of first principles. Accordingly, notwithstanding the celebrity to which he attained, his acuteness in debate, the great prudence of his rules of conduct, his somewhat overstrained conservatism and literal obedience to the laws, and the dignity of his death, he founded no system of philosophy, and Socratism died with Socrates in the prison at Athens. As for the encyclopedic genius of Aristotle, it is only necessary to refer to the titles of his treatises.

Plato, at great expense it is said, bought up the writings of the disciples of Pythagoras, and doubtless his works contain many Pythagorean doctrines. But these are all unquestionably mixed up with and modified by his own views. At any rate (save in Strabo, who notices the geological changes of matter), we are not aware of any classical writer, always excepting Ovid's account of the Pythagorean doctrines, who even alludes to the Phasis of Matter.

Then came the barbaric invasions, and an illiterate and unscientific barbarism. This was followed by a feudalism nearly as barbarous, and a gross and material chivalry. After these, we have the revival of learning. Innumerable facts were discovered regarding the composition of bodies, and the study of chemistry received a great impetus. We had again an age of scepticism and of encyclopedism, and the places of Socrates and of Aristotle were filled, not worthily, by Voltaire and D'Alembert. Happily then came the age of reaction. The decomposition of air and of water, and other important discoveries, led to the true theory of combustion and respiration, to the laws of chemical combination, to the atomic theory; in fact, out of facts sprang principles, scepticism was replaced by faith, and the science of chemistry, with great leading doctrines, was established, and for ever.

Amongst the many great men who have aided in the development of chemical science, it is, perhaps, invidious to make a selection; but certainly our own Davy was not the least distinguished. It is not only that we are compelled to admire the rapidity of his conceptions, the accuracy of his experiments, the clearness of his deductions, and his great discoveries, but Davy is likewise characterised by a unity in his general system of philosophising. Whether we regard him in his laboratory, or by the river side, or at the writing-desk, or preparing for his passage from this world, we perceive the same mind, possessed of a general system; a system indeed imperfect and obscure, as in our present state such must necessarily be, but still not so much a number of isolated opinions as opinions upon different subjects, all having a relation to some great generalisation.

Owing to the labours of this and other great men, the progress of chemistry during this last half century has been amazing, and the beneficial effects that this science has had, and promises even more to have, upon human comfort are very great. One of the discoveries that stands prominent in this respect, was begun about the close of last century by Saussure. He ascertained that vegetables essentially contained a quantity of different kinds of mineral matter, which must, of course, have been obtained from the soil. This discovery of Saussure was not, however, attended to. About thirty years ago, Mr. Grisenthwaite announced that vegetables essentially contained in their structures mineral matters, varying in kind and amount in the different species of plants. He, too, failed in exciting attention. Then Liebig perceived this truth; and owing partly to his fame as a chemist, partly to the energy of his character, and partly owing to his happy knack of creating enthusiastic pupils and followers, it was at once received by the scientific world, and the doctrine of the Phasis of Matter, from earth to plant, from plant to animal, from animal to earth, and so on continually, is now an unquestioned one.

Some, indeed, of the elements contained in the crust of the globe, do not form constituent parts of animal and vegetable structures; but they, too, as will be seen in the following pages, tend to change their relations and combinations with one another.

Thus the word chemistry, as the science has progressed, has gradually enlarged its meaning. At first it simply meant the little knowledge that was possessed regarding gold, silver, &c. And when other metals were better known, the main aim of chemistry or alchemy was to attempt to transmute the inferior metals into gold, and to discover an imaginary powder of life. Even a few years ago, its province was confined to determining the elements and discovering the laws of combination and the compounds that they formed in the inorganic world. But since the publication of Liebig's doctrines, chemistry likewise describes the combination that the elements form in living structures, and the various and rapidly succeeding changes that take place in them. Indeed, the existence of what is now understood by chemistry, depends upon the ascertained fact that the bodies composing this world consist of a variety of elements which, by continually changing their combination, constitute all the substances cognisable to our senses, living or dead. Properly speaking, to this new science a new name should be given; for the laws of combination that prevail amongst the elements in the inorganic world (*i. e.* the laws of the old chemistry), cease the moment these same elements enter a living structure, and other ones take their place. But if this distinction be carefully remembered, and always be clearly kept in view, no inconvenience arises from including under chemistry what more correctly belongs to physiology.

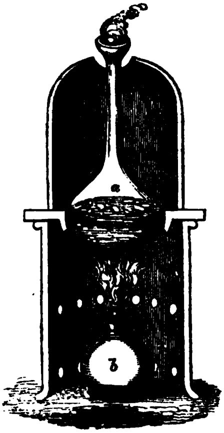
It is to a popular account of this enlarged science of chemistry that the following pages are dedicated. It is intended to describe in them all the important combinations that the elements of matter fall under, whether in the rock, the soil, the water, the air, or in the structures of plants and animals. And as far as the author's ability will permit him, he will endeavour to make the whole a general system of philosophy with regard to matter and its changes. The last part of the book contains an account, necessarily brief and imperfect, of the most important of the existing and probable applications of this enlarged chemistry.

Before actually beginning our task it may, however, be useful and facilitate our future inquiries, if we consider the more general chemical properties of some compound body; water, for example. Water is one of the most common substances in nature. In popular language, indeed, it is called an element; and for long it was regarded by the chemists and natural philosophers as such. It is now, however, as will be immediately mentioned, known to be a compound body.

In the ordinary state in which it occurs in this country, water is a fluid, that is to say, the different particles of any given quantity of it, as for example, a quart, adhere so lightly together that they can glide over one another. If, however, we place two globules or drops of water upon a card very near to one another, they tend very strongly to unite into one. In other words, the particles composing a liquid, although so loosely cohering as to glide over one another, have nevertheless a tendency to cohere together.

If the temperature become colder than that which is commonly experienced in this country, this liquid water becomes solid; that is to say, we cannot move one particle of a lump of ice without moving the whole lump. The force of cohesion is so strong in a solid, that the particles of it cannot glide over one another.

Fig. 1.



If we take some lumps or bits of this solid water or ice, and introduce them into the flask (*a*, *fig.* 1.), and placing the flask in the table furnace *, put under it the lighted oil lamp *b*, we observe that, as the heat is communicated from the lamp to the flask, the cohesion between the particles of the water begins to diminish, and in a little time the contents of the flask are fluid water again. If we go on applying heat thus, water soon attains a temperature of 212°;

---

* For experiments on a small scale, in which it is desired to apply heat, this table lamp furnace is convenient and cheap.

and when this is the case, the cohesion between the particles is destroyed altogether, and the fluid water becomes an invisible vapour, the particles of which have a repelling property to one another, and tend, did other circumstances not present an obstacle, to fly away from one another until they reached the extreme confines of space.

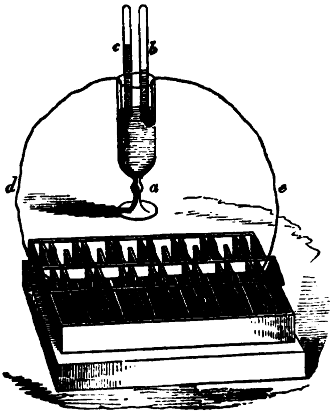
These changes in the relation of the particles are not isolated facts, true only of this substance, water; but all other bodies are subject to a similar law. As a wall is made of so many bricks, so every piece of matter cognisable to our senses is composed of a number of ultimate particles or atoms, so small that they cannot be farther subdivided. So very small indeed are these atoms, that our vision cannot discriminate them, nor detect the joinings of them together. Their distance from one another in any substance varies according to the temperature around them. When the temperature is great, these atoms are farther apart from one another; and when it is little, they are nearer. And in proportion to the distance of the atoms one from another is the power of cohesion, and this is the reason why the same substance, water, is, at different temperatures, a solid, a fluid, and a gas.

But, although all substances are subject to this law, yet the degree of temperature necessary to alter the cohesion of their atoms varies in different substances. Iodine, for example, is solid until it is exposed to a heat of 225°, and does not become a gas until the temperature rises to 347°. Indeed, we have not a sufficient range of temperature in this planet, at least in those parts of it where man can arrive at, to make every substance pass through all these three forms. Alcohol, for instance, becomes a gas at a temperature of 168°; below that, it is a fluid; but no cold that we can produce, can render alcohol solid. Gold, on the other hand, retains its solid form until the temperature becomes so very great at 2016°, and never under any heat we can expose it to, becomes a gas. If the other planets, however, are composed of the same elements as ours is, the case will be different. In the planet Herschel, so very far removed as it is from the sun, not only would alcohol be frozen, but probably the very air itself would be solid; and in the planet Mercury quicksilver would be a gas, and iron and gold would form molten seas.

Another instructive experiment that may be made with water is as follows: — The glass *a*, *fig.* 2. is filled with water; the two test tubes, *b* and *c*, are also filled with water and inverted in the glass. The two wires, *d* and *e*, come from a galvanic battery and enter the two tubes. By and by small bubbles of gas will be seen to form in the tubes, and to occupy the upper portions of them. When a sufficient quantity of gas has been formed, the wires may be withdrawn and the properties of the gas in each tube examined. It will be found that one of the tubes (this depends upon which wire has been in it) contains twice as much gas as the other. If we take the tube marked *b*, containing the lesser quantity, and put into it a glowing shaving, the shaving immediately bursts out into a bright flame, and the gas is in fact oxygen.

If a similarly glowing shaving be put into the gas of the other tube, it is at once put out. The gas contained in it may easily be proved to be hydrogen.

Fig. 2.



This experiment proves that water is not a simple element but a compound formed by the union of oxygen and hydrogen. But let the chemist do what he will, he cannot separate either oxygen or hydrogen into two, and he consequently regards them as elementary substances. When he examines all the bodies found in and around the surface of this earth, he finds that they may all be resolved into some sixty elementary bodies, the various combinations of which form the animals, vegetables, and minerals of the earth; just as the twenty-six letters of the alphabet, by different combinations, form the almost countless number of spoken and written words.

We may carry on this last simile farther. If we investigate a common *fount*, such as is used by a compositor in a printing-office, we find that it is furnished with a great many examples of some letters and with very few of others. There will be found for example 12,000 es and only 200 zs, 9,000 ts and only 400 js, &c. The reason of this is that the es and ts form a great many more combinations than the zs and js. This is just the case with the elementary bodies in nature, and some of them form many more combinations, and exist in larger quantity than the others. By far the most abundant and important are the following:—

Oxygen
Hydrogen
Nitrogen
Carbon
Sulphur
Phosphorus
Chlorine
Silicon
Potassium
Sodium
Calcium
Aluminum
Iron
Lead
Tin
Gold
Silver
Platinum.

We have then learned four important facts:—

First, That the material objects of this world, are composed of different combinations of a few elementary bodies.

Secondly, That each mass of matter is formed of a congeries of atoms, very minute indeed, and not farther divisible.

Thirdly, That the condition of a substance, as to its being solid, liquid, or gaseous, depends upon the relative cohesion of these atoms in it.

Fourthly, That this cohesion can be altered and affected by heat.

It is, as has been so well expressed by Sir Robert Kane, owing to the fact that the bodies that constitute the external world are composed of a variety of elements, united according to certain laws, that the science of chemistry has an existence. Had the globe been composed of iron solely, man indeed could not have had a being, but incorporeal spirits might have dwelt upon it, and such might have studied its motion through space, and deduced the laws of astronomy, or they might have investigated the laws of statics and dynamics, and become natural philosophers; but they never could have been chemists, nor could any idea of chemistry have been formed in their minds.

The object, then, of chemistry is, first, to ascertain which are the elements composing the structure of the earth, and the objects on and around it; secondly, to determine what are the forces that cause them to combine, and what the laws are that rule over their unions and separations; and, lastly, to describe the various properties and peculiarities of the compounds thus formed.

One of the modes which a chemist takes to ascertain the nature of a compound body, is to analyse it, that is, to separate it into its constituent elements. This, as we shall frequently in the following pages have occasion to notice, he contrives to effect by heat, by mixture, or by electricity. In the analysis of water mentioned above, the last of these means is adopted. Chemistry has now attained such a pitch of perfection, that the hand of man can get hold of nothing that the chemist cannot resolve into its elements, and state also the proportions in which these elements exist in it. This applies to all the minerals that form the rocks underneath our feet, to the soils that lie over these, to the various textures of animals and plants, to the atmosphere surrounding us, and to every substance produced by the arts.

In many cases the chemist can do even more than this. He can not only determine what elements a compound consists of, but he can take these elements and from them construct this compound. This is called synthesis. There is, for example, a common mineral called galena, of which more than 80,000 tons are taken annually from British mines. This the chemist, by the art of analysis, pronounced to be composed of the two elements, sulphur and lead. He can also synthetically make this substance, either by mixing sulphur and lead together and fusing them, or by decomposing a solution of a salt of lead by sulphuretted hydrogen, in a manner to be afterwards described. Or if we take two measures of hydrogen gas and one of oxygen, mix them in a tube, and transmit an electric spark through them, these two elements immediately unite and form the compound, water.

But while the chemist can artificially decompose all compound bodies, and resolve them into their elements, he can only synthetically compose one (although a very large) class of compounds. As in studying chemistry it is necessary to have this constantly in view, it is proper, at the very commencement of our inquiries, to acquire clear ideas upon this head.

We have seen that the atoms that compose a piece of any substance, adhere together by virtue of a power which we term cohesion, or mechanical attraction. This cohesion unites together *similar* particles of matter. There is, for instance, the substance so familiar to every one, marble. So great is the power of cohesion between the atoms of marble, that it is with difficulty we can break a bit of it to pieces. Besides this attraction of cohesion, there is another attraction that unites together dissimilar atoms or particles of matter. This is called chemical attraction, or chemical affinity. Each atom of the marble, for instance, is made up of lime and carbonic acid, and these two substances combine together, and remain united together in obedience to this power of chemical affinity.

Now, as long as matter is in the rocks or soil, it forms, in obedience to this power, compounds that chemists can analytically separate into their constituent parts, and which they can synthetically construct by uniting their elements. But, as we have already seen, matter passes from these states (*i. e.* from the inorganic world) into animals and plants. When it does so the elements no longer enter into the same combinations that they previously did, but into other and different ones. Matter in living bodies no longer obeys the laws of chemical affinity, but of vital affinity, the peculiarities of which will likewise fall under our notice in the subsequent pages. The compounds formed under the influence of vital affinity can be analysed, but they cannot be synthetically constructed in the laboratory.

For example, hydrogen and oxygen can be mixed so as to form water; carbon and oxygen, so as to form carbonic acid; sulphur and oxygen, so as to form sulphuric acid; sulphuric acid and magnesia, so as to form Epsom salts; and so on in an endless variety. The starch of plants is composed of carbon, oxygen, and hydrogen; and so is the fat of animals, and there is no great difficulty experienced in separating either starch or fat into these elements. But no art, by mixing these elements, can construct fat or starch. They can only have their origin in a living structure, in which the power of vital affinity is predominant. This peculiarity is even extended to certain purely artificial substances, obtained from undecomposed living structures; as, for example, alcohol.

It is owing to this fact that chemistry is separated into two important divisions; inorganic chemistry, and organic chemistry. The latter teaches the nature and properties of those substances that are formed in animals and plants, with that of the compounds artificially formed from them where there has been no decomposition; while the former investigates those substances which have either never formed part of an animal or plant, or if they have, have passed through the ordeal of putrefaction.

BOOK I.

THE

REVELATIONS OF THE LABORATORY,

OR THE

PRESENT STATE OF INORGANIC CHEMISTRY.

CHAPTER I.

HEAT AND ITS CHEMICAL RELATIONS.

PROPERLY speaking, the laws of heat belong to the province of natural philosophy; but this substance or property exercises so powerful an influence over chemical phenomena, that a knowledge of its more important laws is absolutely necessary to the chemical student. This book, therefore, commences with an outlined account of them.

Heat, or caloric, signifies, in ordinary speech, the sensation felt on touching a hot substance; but, in scientific language, the expressions mean the cause of that sensation. It is generally believed to be a subtle fluid, the particles of which repel one another, but are attracted by other substances. Every material object contains more or less of it, and no substance has ever yet been found altogether deprived of it, for if any, however cold they may be, are placed in an atmosphere still colder, the thermometer indicates the escape of heat from them.

Heat passes from one body to another with great readiness, and indeed all the bodies in the world have a tendency to acquire an equality of temperature, owing to the heat of a warmer body passing on to a colder. Thus, for instance, all the articles in a room very soon become of the same temperature, and if we put into it a heated iron or a jug of hot water, the iron and the water very speedily become no warmer than the furniture, &c., previously in it. It is to this ready transference of heat from a warm to a colder substance that our sensations of heat and cold are owing. When our hand, for example, touches some substance that is colder than the hand, heat passes from the hand to the cold substance and we experience the sensation of cold. Conversely, if we touch a body hotter than our fingers, heat passes from it to them, and the sensation of warmth is felt.

Heat communicated in this last-mentioned manner is said to be communicated by *contact*. But in order to receive warmth from a hot body, it is not necessary to absolutely touch it. When we sit at some distance from the fire, we feel the heat passing from the burning fuel to us. Heat communicated in this manner is said to be by *radiation*, because the heat is supposed to pass from the hot body through an appreciable space in radii or rays. The modes of communication of heat from one substance to another, both by contact and radiation, require a little comment. In considering them, it is first necessary to regard the manner in which a hot body cools, and also the manner in which a cold body warms.

Supposing that two iron balls, equal in size, were suspended from the roof at a little distance from one another, the one cold the other red hot, it would be found in a very short time that they acquired an equal amount of temperature. The heat from the surface of the red-hot ball first passes off by radiation, and then the heat in the interior passes from particle to particle, until it reaches the surface, when it, too, is radiated. The surface of the cold ball first obtains heat, and then passes it from particle to particle to the interior. This passage of heat from particle to particle of the same substance, is called the *conduction* of heat.

It is found that heat passes through different substances with very different degrees of rapidity. If, for example, we put one end of a stick into a fire, we can touch the other end with impunity, even until the whole is nearly burnt away. But if we place one end of a bar of iron in the fire, the other end soon, owing to the heat passing along its particles, becomes so hot, that we cannot bear to touch it. Substances like the latter are called good, and those like the former bad, *conductors* of heat.

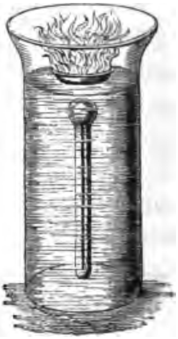
Of all substances the metals are the best conductors of heat, and of metals in this respect the best is gold. The worst conductors of heat among the solids are light porous ones, such as furs, flannels, feathers, and the like. The object of clothing is to prevent the escape of the heat generated by the body, and accordingly some of these bad conductors are always in cold countries, particularly in winter, fixed upon for this purpose. Count Rumford made a very ingenious experiment relative to the conducting powers of the materials commonly employed as raiment. He embedded a thermometer in several of these, and ascertained how long each different substance took to cool through a hundred and thirty-five degrees. He obtained the following result:—

Air	in 576″	Raw silk	in 1284″
Lint	in 1032″	Beaver's fur	in 1296″
Cotton-wool	in 1046″	Eider down	in 1305″
Sheep, ditto	in 1118″	Hare's fur	in 1315″

In the same manner all the substances round about a fire-place ought to be bad conductors; and the reason that massive iron grates give out so little heat into a room, as compared with Kinnaird ones, is that the metal conducts the heat rapidly, and the clay of the Kinnaird slowly. In fact, the conducting power of iron as compared with that of clay is as 370 to 11.

Correctly speaking, fluids scarcely conduct heat at all. If, for example, we invert an air thermometer in a jar of water, with the bulb near the surface, as seen in *fig.* 3. and place a cup containing ether set on fire upon the surface, the combustion may go on for a very long time before any effect is produced upon the thermometer. And even then the heat is probably rather conducted by the glass than the water. It is, however, now ascertained that fluids do conduct heat, although in a very slight degree.

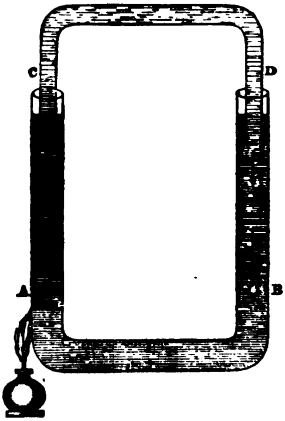
Fig. 3.



Nevertheless, it is a very familiar fact, that if we apply heat to the bottom of a glass containing liquids, the whole mass becomes uniformly hot, and that too very rapidly. This is not, however, owing to the conduction of heat from particle to particle; but owing to each particle when heated being set in motion or to *convection*, as it is sometimes called. When any particle of a fluid is heated, it expands and becomes lighter than those lying above it, which have not received any increase of temperature. It therefore ascends, and a cold and therefore heavier one takes its place at the bottom, and it in its turn becomes heated. The particles also that are ascending part with a portion of their heat to those they actually come in contact with.

The rise of the heated particles of a liquid, and the descent of the cold ones, is illustrated by *fig.* 4. In it A B is a siphon filled with coloured water, and C D another siphon filled with colourless water. If heat be applied at A, the coloured particles will be seen to ascend, and the colourless ones to descend, and thus circulation will continue as long as the heat is applied.

Fig. 4.



As for aëriform or gaseous fluids, they conduct heat even worse than liquids, and the smaller conducting power of bones, skin, furs, &c., mentioned in the table printed a little above, is probably mainly to be ascribed to the great quantity of non-conducting air contained in their porous structures. If, however, heat be applied to aeriform bodies from below, they soon acquire an uniform temperature, exactly in the same manner as fluids do. We have a very familiar instance of this in what takes place in a common fireplace. What we call smoke is a current of heated and ascending air, coloured or rather darkened by soot, and we can often, and in a furnace we can always, hear the roar produced by the stream of cold air rushing in to supply its place. If the fire be very extensive, this rush of air is excessive, and in the memorable burning of Moscow, the currents, or rather hurricanes, thus created by the immense burning mass were so violent, that men and horses passing through the streets could with great difficulty keep their footing.

The two principal conditions that affect the communication of caloric by contact, from one substance to another, are the degree of nearness that the two substances approach to one another and their conducting powers.

With regard to the former of these the rule is, that the more perfect the approximation, the more rapidly is the heat transmitted. Two solids, for example, cannot approach their surfaces so closely to one another as a solid and a liquid, or a liquid and a gas, and this latter still less than a gas and a gas. Hence heat is more rapidly communicated by contact between gases than between liquids, and more rapidly between liquids than between solids.

Respecting the last, we find that when two solids of different temperatures are placed together, the heat from the hotter passes into the colder exactly in accordance with their conducting powers. Thus, if a cubic foot of iron, having a temperature of 300°, and a cubic foot of marble of the same temperature were severally plunged into a gallon of water, the iron would cool much faster than the marble, and the water where the iron was would heat much faster than that in which the marble was, and the reason would be that the conducting power of iron is so much greater than that of marble. The same principle explains the different sensations of heat that we experience when we put our hands upon different bodies that have the same temperature, but different conducting powers. If, for example, we place one hand upon a piece of iron and another upon a piece of marble, each having a temperature of 55°, the iron feels much colder than the marble, for this reason, that the iron more rapidly conveys away heat from our fingers. In like manner, if the iron and the marble had a temperature of 150°, the iron would feel much the hotter of the two, because the heat would be conducted much more rapidly from its interior to the surface next our skin.

The laws that regulate the *radiation* of heat are, to the general student, the most interesting and important of any of the calorific ones; but they do not very particularly concern the chemist, and therefore need not detain us very long. When a hot body is suspended in the air, a part of its heat is *conducted* away by the air, another part is communicated by the particles of air around it, and which ascend on being heated; but the greater part of it is unquestionably radiated. For example, if we suspend a red-hot ball in a vacuum, it loses heat nearly as fast as when suspended in air, although, of course, in a vacuum there can neither be conduction nor convection.

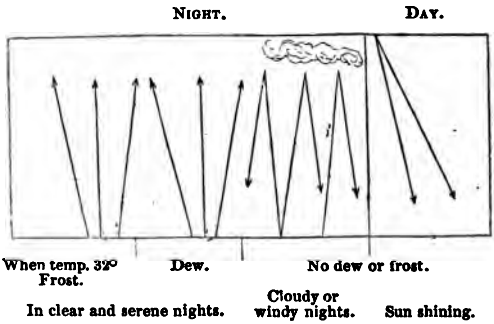
It is a remarkable but well ascertained fact, that the radiation of hot bodies is very much affected by the condition of their surfaces. Thus a bright smooth polished plate of metal scarcely radiates heat at all; whereas if this same surface be made rough, or blackened with soot or the like, it radiates heat well. In like manner a black surface radiates more heat than the same substance having a red surface, and a red surface more than a white one. The former of these statements shows the impropriety of having polished metallic surfaces about fireplaces. On the other hand, if we wish to keep anything hot for a long time, as water, for example, we ought to place it in brightly polished utensils that radiate badly.

When heat radiates from one substance upon another, three things may happen to it; it may be reflected, or it may be transmitted, or it may be absorbed. If, for instance, we place a polished piece of tin sideways before a hot fire, and then hold our hand sideways, but farther from the fire than the tin, we feel the reflected heat impinge upon it. Then most transparent bodies, *i. e.*, bodies that transmit light, also allow of the passage of so much heat through them; but by far the greatest quantity of heat that is radiated from a hot body, is aborbed by other colder bodies; and it is found that the power of absorbing heat is affected very much by the condition of the surface of a substance, and that the conditions of the surface that favour absorption, oppose radiation, and *vice versâ*. Thus, for example, a bright metallic surface scarcely absorbs at all, while a rough and blackened one does so very readily.

It might appear from the foregoing remarks that one set of bodies are always simply radiating heat and doing nothing more, and that another set of bodies are always absorbing heat, and doing nothing more, as far as heat is concerned. But it is now generally believed that every substance is always both radiating and absorbing heat, irrespective of the difference between their relative temperatures. If we adopt this view, we say that the temperature of any substance falls when it radiates more heat than it absorbs, that its temperature remains stationary when the quantity radiated and the quantity absorbed are equal, and that it becomes hotter when it absorbs more than it radiates.

It is only in accordance with this view that we can explain several interesting natural phenomena. We may take as an example of these, the formation of dew and of hoar-frost. Dew falls only when the weather is clear and serene, and it is deposited upon substances that are colder than the air surrounding them. In fact, dew is the water or moisture that did during the day exist in the atmosphere as vapour; but which has become again fluid, owing to its having had its temperature lowered by contact with colder bodies. If, after the dew has been deposited, the temperature fall to or below 32°, it is solidified and frozen, and constitutes hoar-frost. The cause of this transformation of liquid dew into solid hoar-frost is plain enough, and in considering that of the transformation of the atmospherical watery vapour of the day into the liquid dew of the night, the difficulty lies in discovering the cause of the reduction of the temperature. Now, on a clear starlight night, the ground radiates its heat very freely; and on such an occasion, the rays of heat issuing from the earth are cast off into space; no rays are in return radiated from clouds, and, of course, the temperature of the earth's surface falls. If, however, the night be thick and cloudy, the clouds radiate so much of their heat to the earth, and so maintain its temperature. The annexed diagram (*fig.* 5.) will perhaps make this clearer.

Fig. 5.



Then, if we examine the substances upon which dew is deposited, we always find it in the greatest quantities upon good radiators. Thus, on a clear night, grass and leaves can sometimes reduce their temperature by twelve degrees, and these substances are found covered with dew, while bright stones or polished bits of metal, although placed close by them, will be found almost unmoistened.

In the same manner, by increasing the radiating power of water by throwing some loose straw upon its surface, we can often obtain a layer of ice, although the general temperature of the surrounding air does not fall to 32°, the freezing-point. If, however, in such a case, we prevent the radiation by interposing a screen, no ice is formed. It is in this last-mentioned manner that a gardener goes to work when he protects his plants from the bad effects of frost by screening them with matting placed around them. In the same way is to be explained the fact that snow lying over recently sown wheat promotes its vegetation. (We shall hereafter have occasion to see that a certain amount of heat is necessary for germination, and the snow that lies upon the ground, by preventing radiation, promotes the accumulation of it.)

Before leaving this subject of the radiation of heat, the reason of snow lying permanently upon the tops of high mountains may be noticed: one cause, at any rate, that the rays of the sun do not melt the snow at these high mountain elevations, undoubtedly is, that the shining and white snow does not absorb the rays of heat.

The above observations, in a work dealing only with chemistry, may suffice for the laws that regulate the propagation of heat. Some of the effects of this substance (or property) now demand our attention. One of the most remarkable of these is its power of removing the atoms of the substances to which it is applied to a greater distance from one another, or of expanding their bulks.

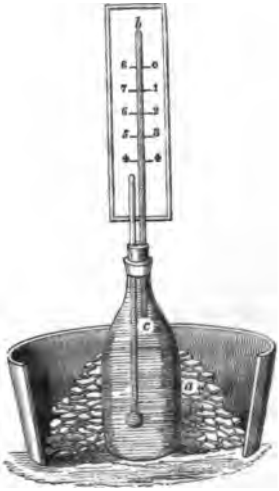
That heat expands solid substances is a familiar fact, and may be easily demonstrated. If, for example, we adapt a ring to an iron rod, so that the ring will just admit the rod, and no more, when cold, we find, if we heat the rod, it will no longer pass through the ring: it (the rod) having become enlarged or expanded. On the other hand, if we heat the ring, it will allow a thicker rod to pass through it: it now having become enlarged or expanded. It is upon this principle that a wheelwright heats to a red heat the iron rim of wheels, and applies them in that state to the wood-work, in order that, as they cool, they may contract (reassume their former size), and firmly bind the wood-work together. The reason that clocks go slower in summer than in winter is that the pendulums are expanded, and therefore lengthened by the summer heat, and consequently vibrate slower. So, also, a nail driven into the wall of an apartment in time becomes loose, for as the weather or atmosphere of the room is hot or cold it expands or contracts, and by doing so gradually enlarges the space that it is in. When, again, we pour hot water into a thick tumbler, the inside next the water sometimes expands so much more rapidly than the outside, that the glass is shivered to pieces. In order to avoid this happening in the laboratory (where the chemist is obliged to be continually applying heat), retorts and other glass vessels intended to be exposed to flame are made very thin.

Different solids have different degrees of expansion from an equal addition of heat. Zinc, perhaps, is the most expansible of all solid bodies. It is scarcely necessary to remark, that when a substance expanded by heat parts with that heat it gradually returns to its original size.

The expansion of liquids by heat is very plainly illustrated by what takes place in a common thermometer. The moment the heat is applied to the mercury we behold plainly enough the expansion taking place. Just as is the case with solids, different liquids have very different degrees of expansion from equal additions of heat. Thus alcohol expands more than water, and water again more than mercury.

There is, however, a most remarkable exception to this law of the expansion on the application of heat in the case of water. Ice, as every one knows, swims upon water, and of course is lighter. It follows of necessity that water in the act of freezing, *i. e.* of parting with heat, expands. This expansion of water may be easily experimentally demonstrated. In *fig.* 6. *a* is a flask, through the cork of which a thermometer is inserted, and also a glass tube, *b*, *c*. If the flask is filled with water to the top, and placed in ice, as the water cools (say from 65°, the common temperature of our rooms), it follows the ordinary law and contracts. It sinks in the tube, and this sinking is indicated in the right-hand scale from 0 to 1, and from that downwards to 4, by which time the thermometer will have fallen to 40°. Then the water begins to expand, as is shewn by the rising

Fig. 6.



of it in the tube, and indicated in the left-hand scale from 4 to 8, and goes on expanding until it reaches 32°, the freezing point. If this frozen water be now thawed the experiment will be reversed, and the ice, warming into water, will contract, that is to say, heat does not expand ice, but, on the contrary, contracts it.

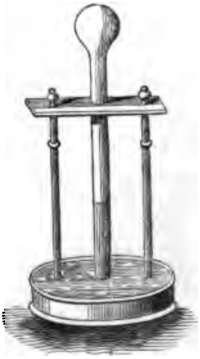
The cause of this is supposed to be that water, upon assuming the solid form, adopts a new and peculiar arrangement amongst its particles. Whatever may be its cause, it is one of the most striking instances of design that can be witnessed in nature, and were it not for it, the globe would scarcely be habitable by man. The freezing of water in winter is of course owing to the coldness of the atmosphere lying above the surface of the water. Therefore, the water at the surface, as it cools, sinks to the bottom; that which was at the bottom rises to the top; and this circulation goes on until the whole bulk of water acquires a temperature of 40°. Then begins this exception to the law of expansion on heating, or contraction, on cooling; and the ice that is formed swims at the top. This ice acts as a covering to the water, and even in very severe frosts it is comparatively thin, and the water underneath is found to have a temperature of 40°. Did water obey the usual law in this respect, the ice would fall to the bottom as fast as formed, and the whole mass of water would become solid.

Gaseous bodies, also, expand upon being heated. If, for example, we partially fill a bladder with air, or any other gaseous matter, fasten the orifice, and hold it before the fire, we behold the expansion. Or if we fill the bladder quite full, and then put it in a hot place, the expansion of the air will burst the bladder. Unlike solids and liquids, gases undergo an equal expansion by the same addition of heat; and it has been ascertained that they all expand $\frac{1}{480}$th of their volume for each degree of Fahrenheit.

It is often extremely desirable to be able to measure the degree of the intensity of heat. This is by means of two instruments — the thermometer and the pyrometer.

The inventor of the thermometer was probably Santorio, a Paduan physician, who constructed his thermiscope (as he called it) in the seventeenth century. *Fig.* 7. represents one of these primitive instruments. It consists of a glass tube, with a ball at one end. The extremity of the tube opposite the ball is open. The air within is rarefied by heating the tube, and the free end is then plunged into some coloured fluid. As the rarefied air contracts, the pressure of the atmosphere forces so much of the coloured fluid up the tube; and when this has happened the instrument is complete. If the temperature of the atmosphere increase, the air in the ball expands, and forces the coloured fluid down; and if the reverse be the case, the air in the ball contracts, and the pressure of the atmosphere then forces the coloured fluid higher up.

Fig. 7.



The objections to this thermometer are, that its dilatations and contractions are so great, that it is not convenient to measure them, and that the movements being influenced by pressure as well as heat, give different indications of the same temperature according to the state for the time being of the barometer.

Towards the middle of the seventeenth century, the members of the Academia del Cimento had the spirit thermometer constructed for their use. Alcohol (coloured) was introduced into a glass tube terminating at the bottom in a ball, from which the air had been expelled by means of heat. The open extremity of the tube was then hermetically sealed, and a scale affixed. As alcohol never freezes, spirit thermometers are very useful for measuring very low temperatures, and are still employed for that purpose; but as spirit boils at an extremely low temperature, the range of temperature that a spirit thermometer can indicate is very small.

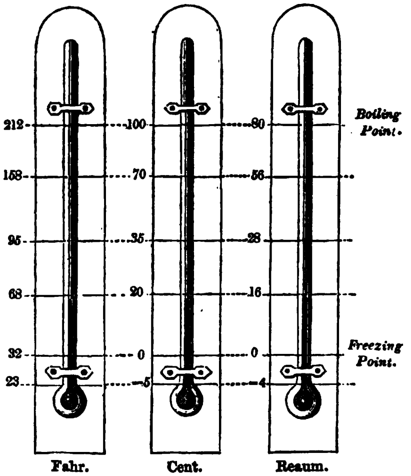
It was this that led Sir Isaac Newton to substitute oil for alcohol. As linseed oil can bear extreme temperatures without either boiling or freezing, this substance would do very well for thermometers, were it not that it adheres so much to the inside of the tube as to be practically useless for this purpose.

Accordingly, the substance now employed for filling thermometers, intended for general purposes, is mercury or quicksilver, a body especially qualified for this use. No fluid can sustain a greater heat without boiling, and none save alcohol and ether can endure a greater cold without freezing. It possesses another remarkable advantage over other fluids and solids, for mercury expands in an increasing ratio, that is to say, the same amount of heat occasions a greater dilatation at a high temperature, than at a low one. But it is found, at any rate, at moderate temperatures, that the ratio of increase of mercury is exactly the same as that of glass, so that in a common mercurial thermometer, although the mercury expands more at a high temperature, than from the same addition of heat at a low one, yet the bore of the tube in the same degree increases its dimensions and so the surface of the mercury does not rise too high.

In order that observations made with different thermometers may be compared together, it is necessary that we have a scale attached to them. For attaining this end the first thing to do is to select two uniform fixed points, and this fortunately is easily managed. When the barometer stands at 30, water boils at an uniform temperature. So also does it congeal at another uniform one. The boiling and the freezing of water thus afford two fixed points always easy to be ascertained. The interval between these two points must next be divided into equal parts. It does not matter how many parts we make, so that every one is willing to adhere to them. Unfortunately, however, this is not the case, and at least three different scales are in use, and their discrepancy sometimes gives a good deal of trouble. These are (see *fig.* 8.) those of Fahrenheit, of Reaumur, and the Centigrade.

The first is the one generally employed in this country. Fahrenheit makes 212° the boiling point, and, strangely enough, he makes 32° above zero the freezing point. There is a tradition that he put 0°

Fig. 8.



where it is from having observed the mercury stand there during a most intense frost in Iceland. Both the Centigrade and Reaumur's thermometer make 0° the freezing point; but the Centigrade divides the interval between this into a hundred, and Reaumur's into eighty parts. The Centigrade is the one usually employed on the continent for scientific purposes, but it is well to know that many French writers upon the arts refer to Reaumur's scale. All thermometers count downwards from the zero, as well as upwards; and in order to distinguish the degrees below zero from those above it, the sign — is prefixed. When we wish to particularly indicate that we refer to the degrees above zero, we write + before the number. But, in general, when no prefix is put, it is understood that degrees above zero are meant.

In order to reduce the degrees of one scale to another, it is only necessary to remember that 9° of Fahrenheit are equal to 5° of the Centigrade and to 4° of Reaumur. The following table shows the degrees of the Centigrade and Reaumur corresponding to those of Fahrenheit, and may perhaps be useful for reference.

Fah.	Cent.	Reaum.	Fah.	Cent.	Reaum.
212	100	80	77	25	20
203	95	76	68	20	16
194	90	72	59	15	12
185	85	68	50	10	8
176	80	64	41	5	4
167	75	60	32	0	0
158	70	56	—	—	—
149	65	52	23	− 5	− 4
140	60	48	14	−10	− 8
131	55	44	5	−15	−12
122	50	40	− 4	−20	−16
113	45	36	−13	−25	−20
104	40	32	−22	−30	−24
95	35	28	−31	−35	−28
85	30	24	−40	−40	−32

Mercury does not boil until it is heated up to 660°*, but so much vapour is given off some degrees below this, that a mercurial thermometer cannot indicate a temperature higher than 630°. If we wish to measure higher degrees of heat than that, we must employ a pyrometer. Many kinds of pyrometers have been invented, but the only one that gives accurate results is that of Daniel.

Daniel's pyrometer consists of two parts — the register A, and the scale B (*fig.* 9.); A is a blacklead case, C an iron rod somewhat shorter, and D an earthenware plug that fits in the case at the end of the iron rod. The iron rod is put into the blacklead case, and the plug is made to abut against it, its other end being against the end of the hand of the index. The instrument is then ready to be submitted to any temperature. The blacklead case scarcely expands at all, but the iron rod is a good deal expanded by heat, consequently the greater the temperature is to which the pyrometer is applied, the more does the rod expand; and the more the rod expands, the further does it push the earthenware plug. The manner in which this is recorded on the index is evident from an examination of the illustration.

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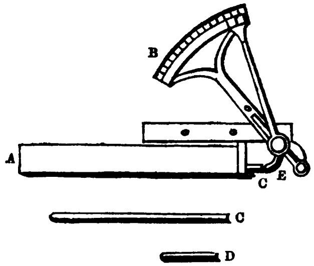
* When not otherwise stated, the degrees of temperature referred to in this work are those of Fahrenheit's scale.

Fig. 9.



The following table (taken from Sir Robert Kane) exhibits the temperature on the scale of Fahrenheit, at which some of the most remarkable effects of heat are produced:

– 135°. The greatest cold ever known by man.

– 121°. The compound formed by alcohol and carbonic acid melts.

−	91°.	The greatest cold produced by a common freezing mixture.
−	58°.	Temperature of planetary space.
−	60°.	The greatest cold observed in the arctic regions
−	47°.	Sulphuric ether freezes.
−	46°.	Nitric ether freezes.
−	39°.	Mercury freezes.
+	1°.	Sulphuric acid freezes.
	14°.	Oil of turpentine freezes.
	20°.	Wine freezes.
	25°.	Blood freezes (*i. e.* blood out of the body).
	32°.	Water freezes or ice melts.
	36°.	Olive oil freezes.
	98°.	Heat of blood in the human body.
	108°.	Phosphorus melts.
	174°.	Alcohol boils (or becomes gaseous).
	201°.	Rose's soluble alloy melts.
	212°.	Water boils.
	218°.	Sulphur boils.
	662°.	Mercury boils.
	810°.	Antimony boils.
	1141°.	Heat of a common fire.
	1869°.	Brass melts.
	1873°.	Silver melts.
	1996°.	Copper melts.
	2200°.	Gold melts.
	2786°.	Cast iron melts.

The indications of the temperature given to us by the thermometer and pyrometer are, both in science and in the arts, of very great importance. Still the amount of information that these valuable instruments afford is very small indeed. They tell us with great exactness the difference of temperature between two bodies, but they give us no indication whatever of the absolute amount of heat that any substance with which they are brought into contact contains. It is clear, for instance, that two thermometers may stand at the same degree in the scale when placed in two different bodies, and yet one of these bodies may contain a hundred thousand times more heat than the other.

For example: if we fill a brewer's vat that contains a hundred thousand gallons with boiling water, and then fill a stone gallon bottle also with boiling water, and place a thermometer in each, both thermometers will indicate a temperature of 212°. And yet the vat unquestionably contains a hundred thousand times as much heat as the jar.

But this is not all: if we take a pound of water, having a temperature of 100°, and mix it with a pound of water having a temperature of 50°, and then introduce a thermometer into the mixture, we find that it has a temperature of 75°; that is to say, the 25° lost by the hotter water have been gained by the cooler. We infer from this that a pound (or an equal measure) of water of any given temperature contains as much heat as another pound (or equal measure) at the same temperature. And this law we find holds good of other bodies; and equal weights of the *same* substance at equal temperatures contain the same quantity of heat. But if we mix different substances together, we find a very different state of matters.

Thus, if we mix a pound of mercury having a temperature of 160°, with a pound of water at 40°, the thermometer in the mixture does not stand at 100°, but only at 45°. But if the mercury be at 40°, and the water at 160°, we immediately find that the mixture has a temperature of 155°. In a similar manner, if a pound of water at 100° be mixed with a pound of spermaceti oil at 40°, the mixture will attain a temperature of 80°; but if the pound of water be at 40°, and that of the oil at 100°, the temperature of the mixture will only be 60°.

In order to witness similar examples, it is by no means necessary that the substances experimented upon should both be fluid. Thus, if we take a pound of copper bar, and heat it in an oil-bath up to 300°, and then immerse it in a pound of water, having a temperature of 50°, both copper and water will attain a temperature of only 72°. In other words while the water has only gained 22°, the copper has lost 228°.

Endless examples of a similar nature might be brought forward to illustrate similar peculiarities. Such were first observed by Black, who deduced from them the law that equal quantities of different substances require unequal quantities of heat to raise them to the same temperature, or to make them equally hot. At first Black thought that the heat contained in any substance was contained within its pores, or the spaces between its constituent particles, and he said that the *capacity* for heat varied in different substances. And, what for some time confirmed him in this view was, that in general the denser a substance was, *i. e.* the less space there was between its particles, the less was its capacity for heat. But further observation and experiments proved that this by no means held universally true, and the term "capacity for heat" is not now used, but the expression "specific heat" employed in its place; and the following theory subsequently proposed by Black is received regarding it.

Heat exists in bodies in two different states. In one it is in chemical combination, exhibits none of its ordinary character, evinces no signs of its presence, but is, as he terms it, latent; in the other it is uncombined and free, and passes readily from one substance to another, affecting in its progress the senses, or to use the expression of Black, it is *free* heat.

Any substance, then, that so combines with heat added to it as to render it latent, has a low specific heat, and *vice versâ*.

The following table exhibits the specific heat of some of the more important bodies :—

Water	1000	Iron	114
Alcohol	660	Copper	95
Nitric acid	442	Silver	57
Magnesia	276	Tin	56
Carbon	241	Mercury	33

The specific heat of substances has probably a very intimate connexion with their molecular and chemical constitution. This, however, is a subject that cannot be entered into here

It will be remarked, that of all substances water has the highest specific heat, and this has been pointed out as a striking illustration of design in the creation of the economy of nature. Owing to it the ocean cannot, even during a tropical summer, be much or rapidly raised in temperature. Almost all the heat that it receives becomes latent. If the sea had the same specific heat as mercury, it would in a single hot day destroy all animal life in it, and very soon be entirely converted into vapour.

This seems the proper place to make the remarks we have to offer regarding the changes produced in the form of bodies by heat.

All bodies are either solid, liquid, or gaseous, and, as before stated, the condition that they are in depends upon the heat they contain. Some substances are always solid, as charcoal; some, as oxygen, are always gaseous; and alcohol is never solid; but it is probable that if we had it in our power to expose these substances to a sufficient degree of heat or cold, we could respectively melt or solidify them.

At any rate, by exposing most substances to different degrees of temperature, we can render them solid, liquid, or gaseous, at will. Thé process of making a solid a fluid, is called liquefaction; and that of making a fluid a gas, vaporisation. Both demand a little notice.

One of the most remarkable of the phenomena that attend liquefaction is the fact that during it a great deal of heat becomes latent. Thus, if we mix together a pound of water with a pound of water at 172°, the mixture will have the medium temperature of 102°. But if we mix a pound of ice at 32° with a pound of water at 172°, the ice dissolves quickly, but the temperature of the mixture is only 32°. The pound of hot water in this latter experiment loses 140° of heat, which enter into the ice, liquefy it, but become quite latent. In other words, when ice liquifies, a quantity of heat becomes latent that would be sufficient to raise a similar weight of water 140°. Or the experiment may be varied: if two flasks, one containing a pound of water at 32°, and another a pound of ice at the same temperature, be placed upon the same fire, by the time the ice has melted, and has only a temperature of 32°, the water in the other pan will have a temperature of 172°.

As so much heat disappears during liquefaction, it follows that when a liquid becomes solid, heat is evolved, or that latent heat becomes sensible or free heat. And such is experimentally found to be the case. Thus water in the act of freezing, and with the surrounding temperature at zero keeps at 32°, and the only possible source of heat is its own latent heat. What is more, if distilled water be kept very still, it may be cooled down to 21° without ice being formed; but the moment it is stirred it freezes, and immediately it does so the temperature jumps up to 32°.

Although, by management, we can cool water below 32° without its freezing, yet, in every case, solid water (and a similar rule holds true of all other solids) melts at that temperature. Hence, in speaking of the scale of the thermometer, to be quite correct we should not say the freezing point of water, but the melting point of ice.

We may here remark, that by adding small quantities of saline matters or vegetable acids to water, we very considerably lower its freezing point. Thus sea-water does not readily freeze, and when it does it is pure water that is crystallised. It is for this reason that the ice islands of the arctic regions when melted are drinkable.

If a solid substance liquify without the application of artificial heat, it abstracts from surrounding objects the heat necessary for its liquefaction. Thus, when we make *eau sucré*, by dissolving sugar in water, the dissolving sugar abstracts heat from the water and renders it cooler. By dissolving other saline substance in water or other liquids, a much greater degree of cold can be produced; and if we use snow or pounded ice, the cold is still greater. It is by acting upon this principle that freezing mixtures are made. These substances were very fully experimented upon by Mr. Walker about half a century ago, and the following tables are extracted from his paper.

FRIGORIFIC MIXTURES WITHOUT SNOW OR ICE.

Mixtures.	Parts.	Thermometer sinks.	Degree of cold produced.
Nitrate of ammonia Water	1 1	from +50° to + 4°	46°
Muriate of ammonia Nitrate of potash Water	5 5 16	from +50° to +10°	40°
Sulphate of soda Dilute nitrous acid	3 2	from +50° to − 3°	53°
Sulphate of soda Muriate of ammonia Nitrate of potash Dilute nitrous acid	6 4 2 4	from +50° to −10°	60°
Sulphate of soda Nitrate of ammonia Dilute nitrous acid	6 5 4	from +50° to −14°	64°
Sulphate of soda Muriatic acid	8 5	from +50° to 0°	50°
Phosphate of soda Nitrate of ammonia Dilute nitrous acid	5 3 4	from 0° to −34°	34°

FRIGORIFIC MIXTURES WITH SNOW OR ICE.

Snow or pounded ice Common salt	2 1	from any temperat. to − 5°	
Snow Common salt Sal ammoniac	5 2 1	do. do. to −12°	
Snow Common salt Sal ammoniac Nitrate of potash	24 10 5 5	do. do. to −18°	
Snow Common salt Nitrate of ammonia	12 5 5	do. do. to −25°	
Snow Concentrated nitrous acid	7 4	from +32° to −30°	62°
Snow Cr. muriate of lime	2 3	from +32° to −50°	82°
Snow Potash	3 4	from +32° to −51°	83°
Snow Diluted nitric acid	3 2	from 0° to −46°	46°
Snow Cr. muriate of lime	1 2	from 0° to −66°	66°
Snow Diluted sulphuric acid	8 10	from −66° to −91°	25°

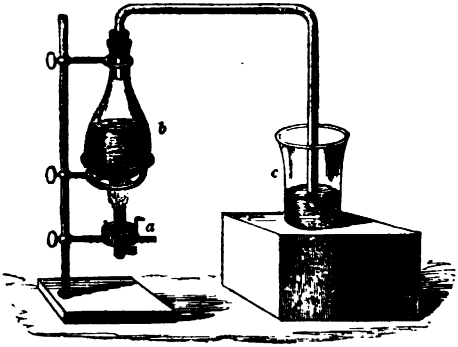
The frigorific power of the above mixtures is increased by previously cooling the materials in a frigorific mixture. It would at first appear that by acting in this manner we could produce any amount of cold, but such is not found to be the case in practice, and the lowest temperature Mr. Walker succeeded in obtaining was — 100°.

As some substances are in the act of being liquefied, a remarkable change comes over their molecular condition. Thus iron or glass become soft and pasty, and two or more pieces can be joined into one. This is a very important and useful property, but one not very easy to explain.

A vapour or a gas (and we may take the terms as synonymous) is produced when most fluids are subjected to heat. Taking water as the type of a fluid, and steam as the type of a vapour, we may say that the former (water) is a compound of ice and heat, and that the latter is a compound of ice and more heat. Sometimes, also, a solid, as sal ammoniac, becomes a gas without previously becoming a liquid, although this is an uncommon occurrence. The passage, too, from the liquid or solid state to that of vapour may take place either slowly or silently, in which case it is said to evaporate; or quickly and violently, when it is said to boil. Evaporation can and does go on at all temperatures, but boiling always begins at a certain temperature, varying for each different substance, and also being affected by the pressure to which it is exposed. We will first notice some of the phenomena common to substances passing from the solid or liquid to the aëriform state either by boiling or evaporation.

One of the most striking of these is, that when a liquid, by boiling or evaporation, becomes gaseous, a great quantity of heat becomes latent. In like manner, when an aëriform body is reduced to the liquid state, a large quantity of latent heat becomes free. Both these statements are easily proved experimentally. We select a practical illustration of the latter. In *fig.* 10. *a* is a lamp suspended under a flask *b*, containing two ounces and a half of water at 32°. Into the neck of this flask there is inserted a perforated cork, in the orifice of which is one end of a bent tube, the other end of which is in the open glass, *c*. This glass also contains two ounces and a half of water, having a temperature of 32°. The lamp in the furnace is then lighted.

Fig. 10.

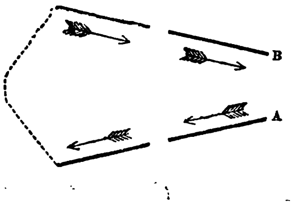


Very soon the water in *b* will begin to form vapour or to boil, and this vapour or steam, as it is commonly called, will pass along the tube into *c*, where it will be condensed. After a little the water in *c* will be observed also to boil, and if the thermometer be introduced into it, that instrument will stand at 212°.

The lamp must now be removed, and the contents of the flask *b* and the glass *c* weighed. Those of *b* will be found to weigh only two ounces, the other half ounce having gone over as steam and been condensed in *c*. The contents of *c* of course weigh three ounces. Yet this half-ounce of steam, that had only a temperature of 212°, has heated to the boiling point two ounces and a half of cold water. The steam has given to the water 932° more heat than itself possessed in a free state; that is to say, so much heat that was latent in the steam, becomes free when this vapour is condensed into water.

The great quantity of heat contained in an insensible form in steam, and the ease with which this vapour can be converted into water or be condensed, and its heat thereby given out in a sensible form, have led, of late years, to its extensive use, both for warming houses and for many processes of chemistry, &c., in which the direct application of flame is objectionable. By means of a series of pipes, steam may be brought from a boiler to every part of a house, and, by being allowed to gradually condense against the inside of the iron pipes conducting it, warm every apartment in it.

Fig. 11.



Before parting with this subject of free and latent heat, we may refer to the adjoined diagram (*fig.* 11.), which exhibits the decrease and increase of free heat in matter during its progress from a solid to a liquid, a liquid to a vapour, a vapour to a liquid, and a liquid back again to a solid. We start with a piece of ice at A, and follow it through out its different stages back again to a piece of ice at B. It is first liquefied, and then evaporated, the result of both being to render heat latent; it is then condensed and afterwards frozen, the result being to liberate or render sensible the heat.

Vapours occupy much greater space than the liquids or solids from which they derive their origin. Thus water passing into vapour expands to 1696 times its previous volume, alcohol to 659 times, and ether to 443 times. Watery vapour is a good deal lighter than the atmosphere (the specific gravity of air being 1000, that of watery vapour is about 620), and watery vapour floats in the atmosphere, forming clouds, and does not descend thence until it is condensed into water.

The temperature at which a fluid becomes converted into a vapour, that rises with sufficient quickness to constitute ebullition, is called its boiling-point, and varies in different substances. Ether boils at 96°, alcohol at 176°, and water at 212°, while mercury does not boil until it attain a temperature of 662°. But in all these cases it is to be understood that the barometer stands at 30 inches. When this is the case, the pressure of the atmosphere is fifteen pounds upon every square inch of surface. To this pressure liquids are of course as much subjected as solids are, and they cannot begin to boil at all until their particles have acquired as much elastic force as will enable them to oppose this pressure of the atmosphere, or they cannot boil until their particles press against those of the atmosphere as strongly as those of the atmosphere do against theirs. But the barometer does not always stand at 30 inches; that is to say, the atmosphere does not always press exactly fifteen pounds upon each square inch of surface. Sometimes it presses rather more, sometimes rather less, and the consequence is that the boiling point of the same substance sometimes varies a little.

For example: in ordinary weather at the level of the sea the pressure of the atmosphere is such that water boils at 212°; but if carried to 1060 feet above the sea's level, the pressure of the atmosphere is so much reduced that it boils at 210°. If the pressure of the atmosphere be altogether taken from it, as is done when some of it is introduced into the exhausted receiver of an air pump, it boils at 72°; and ether similarly situated actually boils at −44°, that is, at a temperature at which mercury would freeze.

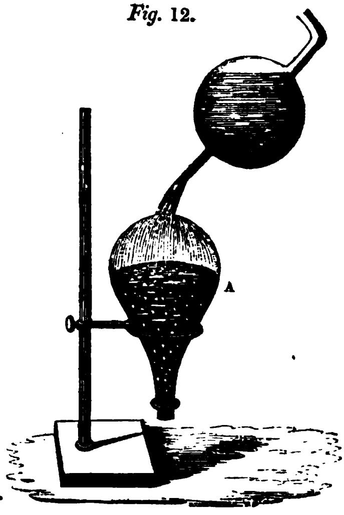
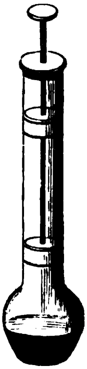


Fig. 12.

In thus boiling liquids *in vacuo* we require to keep the pump working, for if we do not the vapour accumulates and forms an artificial atmosphere, the pressure of which stops the boiling. The fact of the vapour from a boiling liquid accumulating and exercising a pressure on the surface of that liquid sufficient to stop the boiling, is shown by the following little experiment. A, in *fig.* 12., is a flask half-filled with water; this is boiled so as to expel the air, and then carefully corked. It is taken from the lamp and inverted, when the upper part immediately fills with vapour, which by its pressure stops the boiling. If a jet of cold water be now sent upon it, the vapour is condensed, a vacuum is formed, and the water immediately begins to boil again. If, however, warm water be employed, the vapour retains its elastic form, and the boiling stops. We in this case produce boiling by the application of cold, and arrest it by applying heat.

Advantage is taken of the elasticity of steam in the steam-engine, and owing to this, human comfort and civilisation have been most wonderfully extended. The construction of the steam-engine depends upon the much larger space occupied by steam than the water from which it proceeds, and upon the ease with which steam is converted back again to water by cold. Wollaston's cylinder (*fig.* 13.) shows the example of the steam-engine very clearly. It is a glass tube, an inch wide, and blown into a bulb at the bottom. A piston is accurately fitted to the cylinder, the rod of which passes through the brass cap at the top. A little water is introduced

Fig. 13.



into the bulb. If the bulb be held over a spirit-lamp, steam is formed, which immediately drives up the piston. If the bulb be now dipped into cold water, the steam is condensed, and the pressure of the atmosphere immediately forces down the piston again. By alternately applying heat and cold to the bulb, the piston may be moved up and down for any length of time. There is little difference in principle between this little instrument, and the early steam-engine. In that engine, however, the water was condensed in a separate vessel.

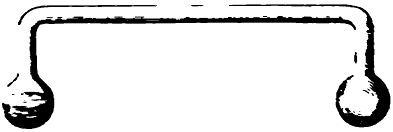
The gentle and quiet mode of vaporisation or evaporation takes place at ordinary temperatures. Thus, if a little water be put into a saucer, it will become less and less, and in a few days it will entirely disappear. All fluids are subject to it, although some evaporate much faster than others, and, generally speaking, those evaporate fastest that have the lowest boiling point.

Evaporation is promoted by a variety of circumstances, such as extent of surface, currents of air, temperature, &c. All these are very evident, and may be passed over.

During evaporation a great quantity of heat becomes latent, and cold is felt. If a drop of ether or of alcohol be allowed to evaporate from the hand, this is easily felt; and if a rag be dipped in fluid and spread over the bulb of a thermometer, we immediately see the mercury fall. By a little contrivance, however, we can even still more strikingly exemplify the great amount of cold that is produced in this manner.

Thus, if a little liquid carbonic acid be poured into the air, a portion of it evaporates with extreme rapidity, and the cold thereby produced is so intense as actually to freeze the other half, and convert it into a white solid. Or, if we place water in an exhausted receiver, and along with it a little saucer containing sulphuric acid, the water evaporates with great rapidity, and the vapour is absorbed by the acid. In this manner we have a perfect vacuum kept up, and the water very soon, by its own evaporation, freezes itself.

Fig. 14.



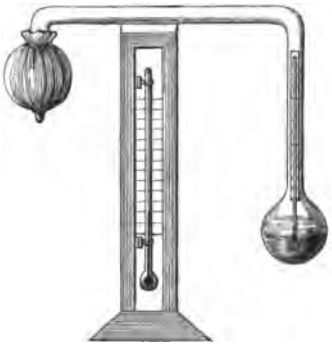
The action of a little instrument contrived by Wollaston, and called the cryophorus, depends upon the same principle. It consists of a long tube, terminating in two bulbs which contain some water, and from which the air has been carefully expelled by boiling (*fig.* 14.). To use it, all the water is brought into one bulb; the remainder of the space, apparently empty, in reality contains watery vapour, the product of evaporation. This watery vapour prevents, by the pressure it exercises upon the surface of the water, any more evaporation from going on. If the empty bulb be now plunged into a freezing mixture, this watery vapour is at once condensed, and then the water in the bulb, which is a yard distant from the freezing mixture, and in nowise affected by it, begins to evaporate so rapidly that the cold thereby generated in the course of about two minutes freezes it.

A certain amount of aqueous vapour is always being sent into the atmosphere by evaporation, from whence it is returned at variable times in the form of rain, snow, &c. There probably is not a drop of water in a spring, a river, or an ocean, that has not once formed part of a cloud; and there is not a cloud of which every particle will not eventually form part of one of our seas. As, however, the evaporation from the seas, &c., of water is greater at one time than another, and as the fall of rain is also variable, it follows that the atmosphere will sometimes contain much more evaporated aqueous vapour than at others. It is often desirable to ascertain the amount that it does so contain, and this is done by means of a hygrometer.

Several kinds of hygrometers have been constructed. Several substances, of which, perhaps, the best example is to be found in human hair, expand when they absorb water, probably owing to their soft texture, and also absorb watery vapour readily from the air. Hygrometers have naturally been for long constructed of such materials, and the hygrometer usually affixed to the barometers of the shops is still one of this kind. But all these materials eventually stretch, and in order to give correct results require to be frequently adjusted to their indices. Another kind of hygrometer was invented by Leslie, and it points out the moisture or the dryness of the air by the rapidity with which evaporation goes on in it. Daniel's hygrometer, the best probably that we possess, is on an entirely different principle from both these.

Warm air can hold suspended in it more watery vapour than cold air. Consequently, when air is cooled by contact with any cold solid body, the watery vapour is deposited in the form of liquid water upon it. Thus, a glass of cold water brought into a warm room has its outside soon covered by moisture derived from the watery vapour of the air. The temperature at which this occurs is called the dew point. If the air be saturated with moisture, the least diminution of temperature is attended with this deposition of dew; while, on the other hand, if the air is dry, there will be a fall of several degrees indicated by the thermometer before this takes place. Daniel's hygrometer, then, consists in determining the dew point. It consists (*fig.* 15.) of a cryophorus, with the stalk, however, of one bulb much longer than that figured in p. 57., and the substance in the interior is not water, but ether. In the bulb with the long stalk, a small thermometer is placed, its bulb being immersed in the ether. The other bulb is covered with muslin. Usually a thermometer is inserted into the stalk, for reference. When the instrument is used the muslin is moistened with ether, the evaporation of which condenses the vapour within the cryophorus, and causes the ether in the bulb to rapidly evaporate.

Fig. 15.



This chills the ether and the bulb containing it, and very soon dew is deposited on the outside surface of the bulb. The height at which the two thermometers stand is then noticed.

Such may, perhaps, serve for an outline of those laws of Heat that more particularly refer to chemistry. Short as has been the outline of them, and imperfect, enough has probably been stated to cause us not to wonder that the mythologists fabled that Prometheus stole fire from heaven; or that the Magi of old worshipped it on the tops of the lofty hills of "the stony girdle;" or that the old Grecian philosophers declared that it must be and was the principle of all things.

CHAP. II.

THE ATOMIC THEORY.

From a very early period in the history of philosophy two opinions have been entertained regarding the ultimate elements of matter. One party maintains that every particle of matter, however small, may be divided into smaller ones, provided our organs were adapted for the purpose. According to the other view, every substance is composed of a number of ultimate molecules which cannot be further divided, and which are therefore called atoms (from ἀ, *not*, and *τέμνειν, to cut*). It is impossible to decide as to which of these opinions is absolutely true, but the probability is in favour of the latter, and the recent progress of chemistry renders it almost certain. To adopt it, moreover, greatly facilitates the acquirement of the elementary principles of chemical science.

As before mentioned, it is possible that the monads of Pythagoras are analogous to the atoms of later writers. The doctrine of atoms was unquestionably taught by Democritus and other cosmological writers of the Eleatic sect, and was still more prominently asserted by Lucretius. It held its ground, too, in mediæval times, and for long after the revival of learning, having indeed been again brought forward and admitted by Newton. It came, however, to be looked upon as a question of no practical consequence; and, indeed, until the discovery of the true elements, it was so.

The experiments of Wenzel, of Higgins, and of Richter, during the latter part of the last century, paved the way to the great and modern atomic theory of Dalton, promulgated fifty years ago. In order to have an intelligible idea of it, it is necessary to have firmly impressed upon our minds the two following facts.

As before stated, all the substances found on or around this world are resolvable into a small number of elementary bodies,—oxygen, hydrogen, carbon, &c. &c.

Any individual compound body invariably consists of the same elements united together in the same proportion. Thus water is a compound of oxygen and hydrogen, and the oxygen, in any given quantity of it, always weighs eight times as much as the hydrogen. No other two elements save these can, by combining together, compose water, nor can these elements do it in any other proportion. We may mix oxygen and carbon together, or hydrogen and carbon together, expose them to any action that we can devise; but by doing this we never will form water. Or, we may put into a tube ninety grains of oxygen, and eight of hydrogen, and send electricity through them. The result will be that eighty grains of the oxygen will unite with eight grains of the hydrogen, and form eighty-eight grains of water, leaving the ten grains of oxygen still uncombined.

We can, however, by mixing (under certain conditions) hydrogen and oxygen, form another liquid besides water, but then the proportions must be different. We must take not only eighty grains of oxygen and ten of hydrogen, but another eighty grains of oxygen, *i. e.*, we must take exactly twice as much oxygen. The compound here alluded to is the binoxide of hydrogen.

Now, Dalton maintained that every simple or elementary substance is made up of a congeries of atoms, the weight of which differs in every one of the different elements. Compound bodies are made up also of ultimate atoms, but every one of their ultimate atoms is composed of the united atoms of their constituents. Of course the weight of each ultimate atom of a compound body is the sum of the weight of its constituents, just upon the same principle as when we mix a pound of sugar with a pound of water we obtain two pounds of syrup.

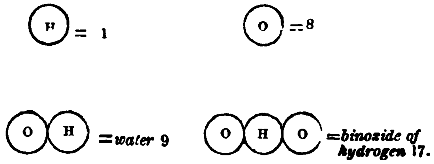
The actual weight of any of these ultimate atoms is so very small that our senses could not observe it, but we may invent a weight inconceivably little. We may name this x weight. If we imagine, then, that there are three simple elementary bodies that we will call A, B, and C, each of which is composed of ultimate atoms, an atom of A weighs 1 x weight, an atom of B 8 x weights, and an atom of C 14 x weights.

Atoms being indivisible things, it is plain that only whole atoms of different elements can join together to form a compound. One atom of A may join with one of B, but one atom of A cannot join with one and a half of B, because there is no such thing as a half-atom; but there is nothing to hinder one of A to combine with two atoms of B, or three, and so on.

If an atom of A, weighing 1x weight combine with one of B, weighing 8x weights, the weight of the atom of the compound so formed will be 9x weights; or if one of A combine with two of B, the weight of the atom of that compound will necessarily be 17.

Now, the weight of an atom of hydrogen is 1, and that of one of oxygen, 8. If these two combine together they form water. If two atoms of oxygen unite with one of hydrogen, they form a peculiar fluid, to which the name of binoxide of hydrogen is given. The weight of the atom of water is of course 9, and of the other substance, 17. The accompanying diagram (*fig.* 16.) will perhaps make this clear.

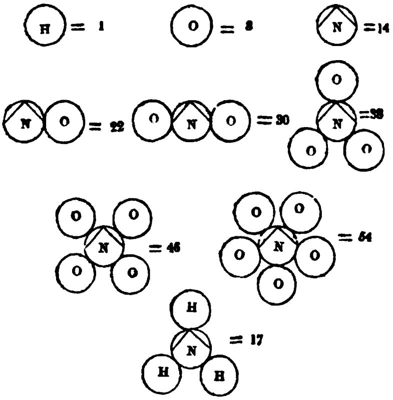
Fig. 16.



Having ascertained that the weight of an atom of B, or oxygen, is x weight 8, if it form compounds with one or more atoms of C, the x weight of which is 14, the x weight of an atom of these compounds will be 14 + 8, or, if there is more than one atom of C, 14 + twice 8, or three-times 8, &c., as the case may be.

The atomic weight of the element nitrogen is 14., and the accompanying diagram (*fig.* 17.) shows the combinations of the atoms of oxygen and nitrogen, and of hydrogen and nitrogen.

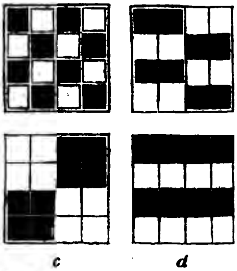
Fig. 17.



Still, notwithstanding the probability of the atomic theory, and the manner in which it facilitates the progress of the student, for many years past almost all systematic writers upon chemistry have explained the laws of chemical combinations in language which is not indeed nearly so clear, but which does not involve the assumption of the existence of atoms and of atomic weights which have not been demonstrated. This fear of hypothesis has doubtless been carried to an extravagant height, and in chemistry, as was said long ago of another science, there are more false facts than false theories. Still, in deference to what has now become usual, the laws of chemical combination will not be considered here, but deferred until the next chapter.

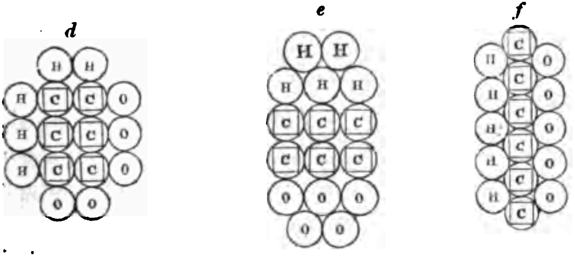
We cannot part, however, with the atomic theory until we have briefly noticed some remarkable facts connected with the composition of bodies which, without its aid, would be perfectly inexplicable. It is clear that the properties of compound bodies will mainly depend upon the kind of the elements that compose them, but it is as clear that they may also depend to a certain extent upon the manner in which the atoms are grouped together. A German writer illustrates this by means of a chess-board, as shown in *fig.* 18.: *a*, *b*, *c*, *d*, are four figures, each composed of eight white and eight black squares. Still, although each has the same number of the same kind of square, yet, owing to the manner in which they are grouped together, we have four different figures. If we suppose that these squares are atoms, the eight black ones, say, eight atoms of oxygen, and the eight white ones eight atoms of hydrogen, then it is evident that these may form four different compounds. Or, in this case, four compounds may exist, composed of the same elements, and in the same proportion, but differing in appearance. In actual practice it is very seldom that such compounds do occur, but they do sometimes, and are called isomeric. Thus sugar, starch, and the principle of wood (lignine) are severally composed of carbon, oxygen, and hydrogen, and they each contain the same proportions of these. Even were each atom of all these composed of one atom each of their constituent elements, we could easily conceive this to be the case. But, in point of fact, each atom of sugar,

Fig. 18.



of starch, and of lignine is a very big atom, and consists of at least six atoms of carbon, five of hydrogen, and four of oxygen. We can readily believe that they are grouped together differently, as seen in *d*, *e*, and *f*, in *fig.* 19.; and we can quite well understand that there should be isomeric bodies.

Fig. 19.



By the aid of the atomic theory we can also understand another interesting, although not so important a fact as the preceding. It is called isomorphism. Besides having a different weight, the atoms of different bodies have different sizes. We will suppose that a compound crystallized body is composed by the union of atoms of *a* with an atom of *b*. We will further suppose that we take away atom *b*, and put an atom of *c* in its place. There may two things happen. The shape and size of the crystallized compound may be very much changed, and this is usually the case; and in such a case we infer that an atom of *c* is of a different size from an atom of *b*. Or the shape and size of the crystal may remain the same. This rarely occurs, but it does sometimes; and then we infer that the atom *b* and atom *c* are of the same size, or are isomorphous. This subject, however, will be recurred to when treating of the properties of alum.

CHAP. III.

CHEMICAL AFFINITY AND THE LAWS OF CHEMICAL COMBINATION.

Chemical attraction, or, as it is now usually called, chemical affinity, is that power which impels an atom of one element to combine with an atom of another element, and form thereby an atom of a compound differing in appearance and properties from either element; or which impels an atom of one compound body to unite with an atom of another compound body, and thus form an atom of a body still more compound. If we hold the atomic theory to be certainly established, the word attraction is the better; but if we think it improper to maintain that, the vaguer expression of affinity becomes the preferable one. In accordance with modern custom the latter is here employed.

Chemical affinity only enables substances to unite and form a compound when the particles of both are in close proximity. In this it widely differs from the mechanical attraction of gravitation. It is quite easy to conceive, however, that this close proximity, which indeed appears to us the closest proximity, is not in all cases as complete as it might be, and that if it were a little more so the affinity might be modified. This, however, is a question very soon to be investigated.

The simplest illustrations of chemical affinity may be obtained by mixing two substances together. Thus, if we take water and alcohol and mix them together, they combine and form grog; and this combination takes place whatever quantities of alcohol and water are chosen. If, however, we mix oil and water together, they do not combine at all. If we mix sugar and water together, they combine in whatever proportions they are in; but if we mix sugar and alcohol together, only a very little of sugar can combine with a great deal of alcohol, and if we put into the spirit more sugar than it can combine with, the portion just remains uncombined sugar exactly as it was ere its particles came into contact with those of the spirit. Conversely, if we put a lump of camphor as big as a marble into a quart of water, the water and a grain or so of the camphor combine, but the greater part of the camphor remains as it was; but if we take a quart of alcohol, we may put a dozen pieces of camphor, each as big as a marble, into it, and they combine with it and disappear immediately.

That is to say, that chemical attraction or affinity is exerted between different bodies with very different degrees of force. Between some bodies it does not exist at all; between others it exists but feebly, and between others, again, it is very strong indeed. A substance *a*, may have no affinity, and form no compound with another substance *b*, while it may have a strong affinity for, and form a compound with a third substance, *c*; and this substance *c*, may or may not have a strong affinity for *b*.

But much more complex phenomena of chemical affinity may be observed than those which attend the simple admixture of two substances, and consequent formation of a third one. If we take, for instance, some ammonia, or hartshorn, as it is popularly called, and mix it with some oil, the two at once combine, and form a white fluid soap or liniment. If to this liniment we add a little sulphuric acid, the ammonia immediately leaves the liniment, unites with the acid, forming sulphate of ammonia, and the oil is set free again just as it was before it united with the ammonia to form the liniment. The reason of this of course is, that strong as may be the affinity of the ammonia for the oil, its affinity for the acid is still stronger.

In this experiment there is not only construction and combination, but destruction and decomposition. Sulphate of ammonia is formed, but the liniment is destroyed. This is said to be an example of simple elective affinity; *elective*, because a substance (in this case, ammonia) manifests a choice for one of two others (in this example for the acid in preference to the oil), by choice uniting with it, to the exclusion of the other. In the course of the following pages we shall have occasion to notice very many illustrations of this simple elective affinity.

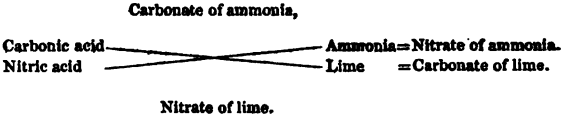
The degree of affinity one substance has for others, or the order in which decomposition will take place, has been expressed by tables. One of these may serve as an instance:—

SULPHURIC ACID.

Baryta
Potassa
Soda
Lime
Ammonia
Magnesia.

This means that sulphuric acid has an amount of affinity for magnesia, a still stronger one for ammonia, one stronger again for lime, a more intense one than that for soda, a greater one yet for potassa, and the strongest of all for baryta. It also means that if sulphuric acid and magnesia are put in close proximity they form a compound — sulphate of magnesia, or Epsom salts. If to a solution of this some ammonia be added, the acid immediately leaves the magnesia (which falls down to the bottom of the glass) and forms sulphate of ammonia. If some lime be added to this compound, the acid immediately leaves the ammonia and jumps to the lime, forming sulphate of lime, or gypsum. However, if we put soda to this, the acid at once leaves the lime and joins the soda, forming sulphate of soda, or Glauber's salts. This in its turn may be decomposed by potassa, and if to sulphate of potassa some baryta be added, the potassa is deserted, and sulphate of baryta formed. The affinity, however, between sulphuric acid and baryta is so strong, that nothing can separate the two. All the substances above any particular substance fixed upon can decompose the compound that sulphuric acid forms with any of the substances printed below.

Much more complicated changes, however, occur than in those decompositions in which we witness examples of the merely elective affinity. In these three substances only are present, and but two affinities exert themselves. There is, for example, sulphate of magnesia, and to this we add some baryta. There is an affinity for the sulphuric acid and the magnesia, but there is another and a stronger between the sulphuric acid and the baryta, and consequently the acid leaves the magnesia and joins with the baryta, forming sulphate of baryta. But it often happens that two compounds, representing four affinities, are mixed together, and *two* decompositions take place, and *two* new compounds are formed. Thus, there are two salts, carbonate of ammonia (consisting of carbonic acid + ammonia), and nitrate of lime (consisting of nitric acid + lime). Their constitution and the changes that take place when they are mixed together will, perhaps, be best seen from the diagram.



It happens that carbonic acid has a stronger affinity for lime than it has for ammonia, and, on the other hand, nitric acid has a stronger affinity for ammonia than it has for lime. When the two salts, then, are mixed together, the acids exchange from the one to the other, and the result is the formation of the two new salts, — nitrate of ammonia and carbonate of lime, or chalk. Occurrences like this (and they are extremely common) are said, in technical language, to be instances of double elective affinity.

Great changes usually accompany these chemical actions in which affinities are brought into play. In general, the substances that combine lose their properties, and the new substance that is formed acquires properties peculiar to itself. There is, for instance, sulphuric acid and potassa, both being very caustic substances. The acid, moreover, is very sour to the taste, and reddens any vegetable blue colour, while the potassa has a very alkaline taste, and makes any vegetable blue that it comes in contact with, green. Yet these two substances unite together and form sulphate of potassa, or sal polychrest, as it used to be sometimes called, which is not caustic, has a bitter taste, and, if brought into contact with vegetable blues, does not affect them at all.

There is also very often a change in density, and two bodies that combine together rarely occupy the same space that their constituents did before they did so combine. Generally speaking, the bulk of the new body is less than the united bulk of the combining substances. The colour, too, is very often affected. Thus, if we combine white mercury and yellow sulphur together, we obtain the bright red compound named vermilion. Iodine has a violet hue, and lead a bluish colour, but the compound formed by these is of a bright yellow colour. Oxide of copper is brown, but most of its salts are either green or blue-white; on the other hand, the salts formed by the yellow oxide of lead are generally colourless.

A change of temperature, also, often accompanies chemical action. When two substances combine together, and occupy less space than they did, of course heat is evolved. So also is it if a gas become liquid, or a liquid solid. If the reverse of this take place, as before explained, then cold is given out. When we come to consider the phenomena of combustion, we shall have occasion to notice another manner in which chemical action causes an evolution of heat.

Another very common change that goes along with chemical action is alteration of form. Solids become liquid, liquids solids, and both sometimes gaseous. The decomposition of gunpowder by heat is a common example of a solid becoming gaseous, and in the following pages we shall often have occasion to notice similar instances.

It is very important always to bear in mind that chemical affinities are liable to be modified, and are sometimes very much modified by certain conditions. As before mentioned, vitality is one of these; and chemical elements forming part of the structure of an animal or a plant, form altogether different compounds from those which the same elements do, as long as they are in the laboratory of the chemist, or in any manner, in the inorganic world. Carbon, hydrogen, oxygen, and nitrogen, if existing together in the non-vital world, form carbonic acid and ammonia; but if they exist together in the structure of an animal, they tend to form flesh. But when the animal dies, this flesh is actually converted into carbonic acid and ammonia. And even in the non-organized world, we can clearly see conditions present that modify and are capable of counteracting or promoting the tendency of chemical affinity. It will be useful to briefly consider some of these.

One of the most important of these is cohesion. Sulphur and iron have, for example, an affinity for one another. But if we place a lump of sulphur by the side of a lump of iron, they do not unite, and this clearly owing to the state of cohesion in which they are, and which prevents that close proximity of them which, as we have before stated, is essential to chemical action. But by dissolving the sulphur by means of heat, and then adding the iron in very small laminæ, we can form sulphuret of iron, or pyrites. In this experiment we diminish the cohesion of the iron by mechanical means, and that of the sulphur by changing it by the application of heat to a liquid state. Both these two plans are often employed in order to promote chemical unions, but the latter is by far the more effectual, and the more frequently used. We perhaps cannot have a better illustration of the power which heat has of assisting the tendency of chemical attraction than in what takes place in many instances of crystallization.

If, for example, we take a salt, say alum, and drop it into water, the alum and the water unite together up to a certain point. When this point, called saturation, has been arrived at, the water can dissolve or combine with no more alum. In practice it is found that eighteen ounces of cold water can dissolve about one ounce of alum. If we put this quantity of a saturated solution of alum in an open vessel, some of the water would gradually evaporate, and in time a quantity, say an ounce, would have disappeared, and the consequence would be, that the seventeen ounces left behind could not dissolve the whole ounce of alum; and the quantity of it that the ounce of water that has disappeared held in solution, say thirty grains, is separated, and forms in large crystals at the bottom of the dish. If, however, we boil the water before putting in the alum, the boiling water communicates part of its heat to the alum, and thereby so increases the combining power of the alum that we may put in actually eighteen ounces of alum before the mixture becomes a saturated solution. As this solution, however, gradually cools, the affinity between the two solutions becomes less and less, and by the time it has quite cooled, the whole additional seventeen ounces of alum will be found uncombined with water, and crystallized at the bottom of the dish.

Cohesion also affects affinity in a manner that it will be necessary here only to state. When we mix together three substances, and if one of the two compounds that may be formed is insoluble, that insoluble one is always formed.

Contrary to what at first would be expected, elasticity, so far from promoting, has actually an unfavourable effect upon chemical action. Two gases placed together rarely unite. The cause of this is that their atoms are too far removed and not in sufficiently close approximation for the play of chemical affinities. It is owing to this that many compound substances may actually be decomposed by an exposure to a high temperature. The expansion in such a case removes the elements of the compound to such a distance from one another that the chemical attraction loses its influence. This is particularly the case when one of the elements combined has a tendency to be volatile, and when the other is fixed. Thus, if the oxides of several of the metals be exposed to a high temperature, the oxygen flies off as a gas, and the pure metal is left behind.

It will be observed, however, that although heat sometimes favours and sometimes actually impedes chemical affinity, yet that the rationale of its action is always the same. When the cohesive power tends to prevent combination, the heat diminishes the cohesive power, or even converts a solid into a liquid, and hence in such a case it favours the action of affinity. But when it renders a substance gaseous, it removes their particles to too great a distance.

Quantity of matter has also a very considerable effect upon chemical affinity. Supposing a substance, say lead, unites with another, oxygen, in more than one proportion, the compound that contains the least quantity of oxygen will be the most difficult to decompose. Thus the peroxide of lead yields its oxygen upon the application of a moderate amount of heat; the red oxide requires a much greater heat; and let us heat the protoxide as we will, we cannot separate the oxygen and the lead.

It is quite possible to express the proportions in which bodies unite, and the laws of combination, without reference to the atomic theory. Four great laws exist with regard to combining proportions, which may be stated as the laws of—1. Constant Proportion; 2. Reciprocal Proportion; 3. Multiple Proportion; and 4. Compound Proportion. Each of these requires a little enlargement and explanation.

1. Constant Proportion. The composition of all chemical compounds is fixed and invariable, and the elements that compose any compound are in it in the same proportion by weight. There is common salt, for instance, which invariably consists of 36 parts of

chlorine to 24 of sodium, and whether we obtain the salt from a salt mine at home, or from our seas, or obtain it from the Pacific Ocean, or artificially compose it in the laboratory, its constitution is exactly the same, and its chlorine and sodium are precisely in the same proportion.

The importance of this law in those arts that have for their end the formation of chemical compounds is very great. Glass, for instance, is a compound of silica and potassa, united together under the influence of heat. Now, if the silica and potassa be put into the furnace in the proper proportions, there is no waste; but if there be too much potassa (and potassa is a pretty dear substance) introduced, that excess is simply wasted. In like manner, in soap-making, and many other manufacturing processes, due economy can only be satisfactorily obtained by the aid of chemical science.

2. Reciprocal Proportion. This is supplementary to the preceding, and teaches that not only does the same compound always contain the same proportion of the same ingredients, but that the proportion in which an ingredient or element enters into one compound, is also the proportion (with a modification, as it were, to be explained in the next paragraph) in which it occurs in *all* compounds whatever. Thus water is composed of oxygen and hydrogen, the oxygen being in the latter in the proportion of 8 to 1. Whenever either of these elements unites with other bodies the same proportion is observed. Thus both unite with carbon (which we may represent by 6) to form carbonic acid and olefiant gas, respectively, and the

carbonic acid is 6 carbon + 8 oxygen, and the other gas 6 carbon + 1 hydrogen. Those who object to the atomic theory as too hypothetical name these combining weights of elements their equivalents.

3. Multiple Proportion. One element often combines with another in more than one proportion, but in such a case the higher proportions are multiples of the lowest. Thus oxygen combines with nitrogen in 5 proportions. One equivalent of the latter combines with 1 equivalent or 8 of oxygen, also with 2 or 16, with 3 or 24, with 4 or 32, and with 5 or 40; but never with any intermediate proportions.

4. Compound Proportion. This is that the combining proportion of a compound body is the sum of the combining proportions of its constituents. Thus water consists of oxygen 8 + hydrogen 1, and its combining proportion is of course 9; carbonic acid consists of carbon 6 + oxygen 8, and its combining proportion is of course 14; and so on for all compound bodies.

CHAP. IV.

CHEMICAL NOMENCLATURE AND NOTATION.

The science of chemistry has progressed so rapidly that chemical nomenclature is not, as it was a few years ago, almost perfect. Indeed, many of the advances that chemistry has made are to be attributed to the accuracy created in the minds of those studying it by the exactitude of its nomenclature; and even yet it excels in this respect all other sciences.

It was only, however, towards the close of the eighteenth century that the present system of naming chemical compounds was adopted. Previously, either arbitrary or fanciful names were applied. The alchemists bestowed upon their chemicals quaint but sometimes expressive names; as, flowers of sulphur, spirit of salt, aqua regia (from its dissolving gold), &c. Astrology also gave appellations to chemicals. The metal quicksilver, or hydrargyrum, is to this day almost always called Mercury, because the astrologers considered it under the influence of that planet. Iron, they held, was under the control of Mars; and physicians have only lately ceased from calling oxide of iron in their prescriptions crocus Martis. In like manner lead, they said, was subject to Saturn, and to this day a solution of acetate of that metal is known as

Saturnine lotion. Afterwards an erroneous opinion regarding phlogiston conferred names upon many substances. It was, however, the discovery of oxygen gas that led to an improved nomenclature, and Lavoisier and his compeers proposed the system now used. This, with the additions that have since been made to it, is as follows.

The names of the elementary bodies are quite arbitrary. Some have existed from time immemorial, and are vernacular, as with us are common names; but all those that have been discovered since the time of Lavoisier have been named from some striking property. Thus, chlorine is so called from its green-yellow colour (χλωρὸς), iodine from its violet colour (ἰώδης), &c. The following is a list of the elements, leaving out one or two very unimportant ones. The symbol that follows them will be explained immediately.

Hydrogen	H.	Strontium	Sr.	Tungsten (wolfram)	W.
Oxygen	O.	Calcium	Ca.	Molybdenum	Mo.
Chlorine	Cl.	Magnesium	Mg.	Tantalum	Ta.
Bromine	Br.	Aluminum	Al.	Chromium	Cr.
Iodine	I.	Glucinum	G.	Vanadium	V.
Fluorine	F.	Zirconium	Zr.	Uranium	U.
Sulphur	S.	Thorium	Th.	Aurum (gold)	Au.
Selenium	Se.	Tellurium	T.	Iridium	Ir.
Phosphorus	P.	Cerium	Ce.	Osmium	Os.
Nitrogen	N.	Lanthanium	Ln.	Platinum	Pl.
Boron	B.	Manganesium	Mn.	Stannum (tin)	St.
Silicon	Si.	Ferrum (iron)	Fe.	Plumbum (lead)	Pl.
Carbon	C.	Cuprum (copper)	Cu.	Bismuthium	Bi.
Kalium (potassium)	K.	Titanium	Ti.	Argentum (silver)	Ag.
Natron (sodium)	Na.	Arsenicum	Ar.	Palladium	Pd.
Lithium	Li.	Stibium (antimony)	St.	Rhodium	R.
Barium	Ba.				

It will be observed that the above names are more Latin than English, and that the symbols are invariably taken from the Latin name. The reason alleged for this is that it insures a universal chemical language. It is only when employing symbols, however, that this rule is rigidly adhered to, and when the words are written at full length the vernacular is used. Thus, iron, lead, and gold, are so termed, and not ferrum, plumbum, and aurum.

The metallic elements all end in *um*, and the only non-metallic body that has this termination is selenium. The three elements that end in *on* —boron, silicon, and carbon — are analogous in their chemical relation. So also are the four that end in *ine* — chlorine, iodine, bromine, fluorine — in theirs. No community of chemical properties is meant by the termination *gen;* and sulphur and phosphorus are names given to these elements before the present system was called into existence. This also is the case with many of the metals, and the whole of these bodies are named quite arbitrarily.

The names of compound bodies are not arbitrary at all, and are intended to indicate not only the class to which they belong, but their composition also.

All these compound bodies may be arranged into three classes. 1st, that of binary compounds, the members of which are composed of two elements only, as water, composed of oxygen and hydrogen; sulphuric acid, of sulphur and oxygen; or common salt, of chlorine and sodium: 2nd, that of ternary compounds, the members of which consist of three elements, as chalk, which is composed of carbon, oxygen, and calcium: and, 3rd, that of quaternary compounds, the members of which essentially consist of four elements.

Among the binary compounds a foremost position must be given to the unions formed by the elementary bodies and oxygen. Of these, two great divisions exist: those which are acids, i. e. redden vegetable blues, taste sour, &c.; and those which do not possess these properties, and which are denominated oxides.

Many elementary bodies unite with oxygen in more than one proportion. When this is the case, the different oxides are named as follows: the oxide containing one atom of oxygen with one atom of the other elementary substance that combines with it, is called the protoxide (from *πρωτός*, *first*, or *head*); if, as is sometimes the case, there is a union formed in which two atoms of the element unite with one of oxygen, this compound is termed a suboxide; if one atom of the element unite with two of oxygen, the compound is the deutoxide; if with three, the tritoxide; and if with four, the quadroxide.

Deutoxide simply means the second oxide. If this deutoxide contain twice as much oxygen as the protoxide, it is often called the binoxide, from the Latin *bis*, twice. And sometimes the compound that contains the greatest quantity of oxygen is named the *per*oxide, from *per*, which signifies superiority or perfection.

We may illustrate the above. One atom of copper unites with one of oxygen, and forms the black oxide of copper. This is named the protoxide. Two atoms of copper also unite with one of oxygen, and form another compound called the red oxide of copper. This is, of course, the suboxide of the chemists. One of the oxides of iron, again, contains two atoms of iron united with three of oxygen*, and as this is the oxide of iron richest in oxygen, we may call it the peroxide.

If the compound formed by the union of any elementary body with oxygen be acid, the word acid is added to that of the other body, to which also *ic* is appended. Thus, sulphuric acid means a compound of sulphur with oxygen that possesses acid properties. Sometimes, however, the same element unites in different proportions with oxygen, forming different compounds, all of which are, moreover, acid. In this case, that which contains less oxygen than the acid that ends in *ic* is called the acid ending in *ous*. For example: sulphur forms an acid with oxygen that contains less of the latter than sulphuric acid does, and this is therefore named the sulphurous acid. A few elements, however, as is the case with sulphur, form with oxygen more than two acids. In such cases the Greek prefixes, *ὑπὸ*, *hypo*, and *ὑπὲρ*, *hyper* (*under* and *over*), are employed and prefixed to the names just mentioned. Thus, there is an acid of sulphur containing less oxygen than the sulphurous, and this is called the hyposulphurous acid; and there is another that contains more oxygen than the sulphurous, although less than the sulphuric, and this is named the hypersulphurous acid, or the hyposulphuric.

Sometimes, too, the acid that contains the greatest quantity of oxygen is named the per-acid: thus, chlorine forms with with oxygen four acids—the hypochlorous acid, the chlorous, the chloric, and the hyperchloric or perchloric acid.

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* See the oxides of iron enumerated at p. 196.

Binary compounds of the following substances are named as follows: —

Those of	Chlorine,	form	Chlorides.
	Iodine,	„	Iodides.
	Bromine,	„	Bromides.
	Fluorine,	„	Fluorides.
	Sulphur,	„	Sulphurets.
	Carbon,	„	Carburets.
	Phosphorus,	„	Phosphurets, &c.

In order to distinguish between different chlorides, sulphurets, &c. of the same element, Greek prefixes are used, as in the case of the oxides. Thus we have protosulphurets and deutosulphurets, and if any of the latter contain twice as much of the sulphur, &c., as the former, it is often called the bi-sulphuret, and so on.

When a non-metallic element is united with a metallic one, the compound is always named, as above, iodide of sulphur, sulphuret of iron, and the like; but when two non-metallic elements unite together, it is somewhat difficult to tell which name to put first. Chlorine and phosphorus unite together, and it may be doubted whether to call it chloride of phosphorus or phosphuret of chlorine. It is of very little consequence which of the titles we bestow upon it, but the one is put first which the chemist thinks upon the whole stands highest in the scale.

The compounds of arsenic and tellurium are named arseniurets and tellurets, because these metals in their chemical relations very much resemble sulphur, phosphorus, and other non-metallic elements. But the compounds that the other metals form with one another, or alloys, as they are called, receive no special chemical name save that if they contain mercury as a constituent they are styled amalgams.

When Lavoisier prepared his system of nomenclature, he believed that all acids essentially contained oxygen. This, however, is now known to be incorrect. Hydrogen and sulphur, for example, form binary compounds that are acid. When this is the case, the prefix *hydro* or *sulpho* comes before the name of the oxacid. Thus, sulphuric acid means an acid compound of sulphur and oxygen, while hydrosulphuric acid implies that acid compound that contains no oxygen, but which consists of sulphur and hydrogen.

The union of two primary compounds forms a secondary one. The names most frequently employed regarding these are acid, base, and salt. A base is a substance that, combining with an acid, forms a compound of which it is, as it were, the basis. The compounds thus formed receive the name of salt, because in external character they resemble common salt. The term salt is now, however, extended to many binary compounds, as to iodides, chlorides, &c.

It will be observed that an oxacid never combines with a simple body, but always with a compound one, and which is invariably an oxide. Thus, sulphuric acid (which is composed of sulphur and oxygen) can form no union with magnesium; but with magnesia (which is composed of magnesium and oxygen) it can, and does, forming sulphate of magnesia, or Epsom salts.

As, however, the acid and the base, the union of which forms oxysalts, have a common constituent, namely, oxygen, these salts are made up of three elements only. They are always named, however, as if the acid was, although combined with, yet distinct from, the base.

The nomenclature of salts is very simple, and at the same time expressive. If they contain an acid that terminates in *ic*, the name of the salt ends in *ate*. Thus, all salts formed by sulphuric acids are called sulphates; by phosphoric acid, phosphates; by nitric acid, nitrates; and so on. If the acid ends in *ous*, then the salt terminates in *ite*. In this manner, the salts formed by nitrous, phosphorous, and sulphurous acids, are denominated nitrites, phosphites, and sulphites.

It is common, for the sake of brevity, to say sulphate of iron, nitrate of copper, &c. Such expressions, however, are quite erroneous, as no such compounds exist. It ought to be sulphate of oxide of iron, nitrate of oxide of copper. However, certain oxides were familiarly known and received names before it was known that they were oxides. The oxide of sodium was, for instance, called soda; and that of calcium, lime. It is quite allowable to continue the use of these names, and, in place of saying sulphate of oxide of calcium, to abbreviate it into sulphate of lime.

The same acid often combines in more than one proportion with a base. In this case numerical prefixes are employed, in the same manner as in the instance of the binary compounds. Thus, there is the carbonate of soda, every atom of which contains one atom of carbonic acid and one of soda; and also the bicarbonate, every atom of which contains two atoms of carbonic acid. Sometimes, again, the base is in excess, and then the Greek numerals *dis*, *tri*, &c.,

are prefixed. Thus, there is the disulphate of oxide of zinc, every atom of which has two of base and one of acid. Of late, it has become customary to call these compounds bibasic, tribasic, &c. The salt just referred to would, according to this nomenclature, be the bibasic sulphate of zinc.*

Acids containing hydrogen, &c., instead of oxygen, have their last syllable changed in the same manner. The nomenclature of the salts that they form, together with some other additional laws of chemical nomenclature, will be noticed afterwards.

Quaternary compounds generally consist of two salts united together, each containing the same acid. Thus, common alum is a compound of sulphate of alumina and sulphate of potassa. In such a case the acid is only expressed once, and this substance is called sulphate of alumina and potassa. Sometimes, however, more complicated quaternary compounds are formed, and then there is nothing for it but expressing their names at their very full length. Such compounds, however, are rare and unimportant; and even were they not so, no practical inconvenience is felt, owing to the system now universally adopted of chemical notation.

In the list of elementary substances at p. 81. a letter, or two letters, will be observed to follow each name. When one letter only occurs, it is the initial of the name; and when there are two, the second one is either the second letter of the word, or some other one of it. These are called symbols, and are used not only to denominate the name of a substance, but also its atomic weight. Thus, O not only means oxygen, but one atom of it. If we wish to express three atoms of oxygen or five atoms of nitrogen, we write $3O$ and $5N$.

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* For a further explanation of the meaning of these expressions see pp. 142, 143, &c.

When it is wished to express compounds, the symbols of their elements are written with + between them. Water, for instance, is composed of one atom of oxygen united to one of hydrogen, and is therefore written $H+O$. Nitric acid is a compound, each atom of which consists of one atom of nitrogen united with five atoms of oxygen, and it is therefore written $N+5O$. Potassa is composed of one atom of potassium and one of oxygen, and is therefore written $K+O$. Nitrate of potassa consists of one atom of potassa and one of nitric acid. The following formula would then indicate its composition: — $K+N+6O$. But this does not tell us how much oxygen is united with the potassium, and how much with the nitrogen, and we therefore enclose in brackets the symbols of those elements that are supposed to be united, and unite them thus: $(K+O)+(N+5O)$. A number placed on the outside of a bracket indicates the number of atoms of the substance contained within it. For example: $(K+O)+(S+3O)$ is sulphate of potassa, and $(K+O)+2(S+3O)$ is bisulphate of potassa.

Sometimes, in the case of binary compounds, the + is scarcely necessary, and water, for instance, may be written either $H+O$, or HO.

In writing the symbols of binary compounds the substance least resembling oxygen is usually written first.

Sulphuric acid, for example, is not written $3O + S$, but $S + 3O$. In like manner, in the formulæ of oxysalts the symbol of the base always goes first,—sulphate of potassa not being written $(S + 3O) + (O + K)$, but $(K + O) + (S + 3O)$. The reason of this regulation is, that the symbols express not English but Latin words, and the construction of the Latin language requires that the genitive precedes the nominative that governs it.

The above very simple system of notation is the one that will be employed in the following pages. It is proper, however, to mention some modifications of it now often employed.

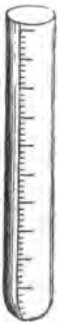
It is common to dispense with brackets, and to denote the number of atoms of an element by a figure following it. In this manner, nitrate of potassa is written $KO, NO5$. The degree of oxidation is often expressed by dots over the symbol, every dot representing an atom of oxygen. $\dot{H}$ is water, $\dddot{S}$ sulphuric acid, according to this plan. Sulphur is sometimes represented in the same manner, a comma being substituted for the dot. Other modifications too are used, but may here be omitted.

CHAP. V.

LAW OF GASEOUS VOLUMES.

It was proved by Gay Lussac that there exists a relation between the atomic weight of gases (or vapours) and their volumes. This is of considerable practical importance in the laboratory, as it is much easier to measure than to weigh a quantity of gas. The usual manner of collecting gases is in round glass vessels, similar to *fig.* 20. marked with lines indicating how many cubic inches of gas the vessel, or different portions of it, can contain. This is filled with water or mercury, as the case may require, and inverted into a pneumatic trough, also filled with water or mercury. The gas is then passed up through the water or mercury, and displaces it, or a portion of it. The number of cubic inches of gas is then easily ascertained by looking at the mark at the line where the gas and water meet.

Fig. 20.



Gases expand and contract according to the temperature and the pressure of the atmosphere. In experimenting upon them it is therefore necessary that both these be at some fixed and uniform rate. It is usual to select the temperature of 60°, and a pressure of the atmosphere indicated by the mercury in the barometer standing at 30 inches.

If we take eight grains of oxygen (8 is the atomic weight of oxygen), and introduce it into the graduated jar, we can easily ascertain how many cubic inches it occupies. If we then take thirty-two grains of phosphorus (32 is the atomic weight of phosphorus), and reduce it to vapour, we find that it occupies exactly the same number of cubic inches as the eight of oxygen do. Or, if we take seventy-five grains of arsenic (75 is the atomic weight of arsenic), and vaporise it, it also is found to fill precisely the same space as either eight grains of oxygen or thirty-two of phosphorus.

As far as regards the volume of an atom of gas, oxygen is taken as the standard, and its property in this respect is denoted by a symbol consisting of a square, □. Against the names of all other substances that agree in this respect with oxygen we make a similar symbol, and thus save unnecessary repetitions of the fact.

Some substances, however, do not yield volumes identical with that of oxygen. If we take one grain of hydrogen (1 is the atomic weight of hydrogen), it is found to occupy twice as much space as the eight grains of oxygen. Indeed, this was noticed to be the case in the decomposition of water noticed in the Introduction. Water is composed of an atom of oxygen and an atom of hydrogen, but when the two are separated the hydrogen occupies twice the space. This is because its specific volume is to that of oxygen as 2 to 1. Substances agreeing in this respect with hydrogen have this property represented by the symbol □□.

All the elementary bodies give a volume equal to that of either oxygen or hydrogen. Some of the compound ones, however, give volumes four times that of oxygen, and this is denoted by four squares ⊞.

The practical application of this very curious law will now be evident. If we wish to form a compound by combining its elements that exist in a gaseous form, we do not weigh them, but ascertain their specific volume, and measure them accordingly. If, for example, we wish to synthetically form water, we take one volume of oxygen and two of hydrogen, and transmit the electric spark through the mixture.

When gases combine with one another the volume of the compound is sometimes the sum of the volume of the gases forming it; but in general there is condensation, and the volume of the compound is less than that. Thus, two volumes of hydrogen and one of oxygen form only two volumes of steam. The condensations that occur, however, are very simple ones, and the combining measures of compound gases are always either the same as those of elementary bodies or multiples of them.

CHAP. VI.

THE NON-METALLIC ELEMENTS AND THEIR BINARY COMBINATIONS.

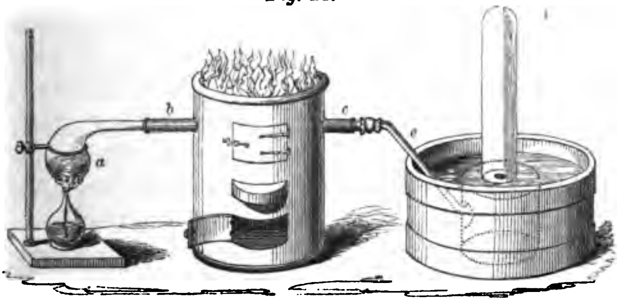
I. Hydrogen.

From an early period in the history of chemical science, gases were known that were characterized by being inflammable on contact with flame. Afterwards, two kinds of inflammable gases were recognised — heavy and light. This latter was then named phlogiston, from the opinion that it was the essence of heat. It was first minutely examined by Cavendish, nearly a century ago, and pronounced an elementary substance. It received its present name, ὕδωρ, *water*, and γεννάω, *to generate*, from the French school, on account of its being an essential constituent of water.

It never in nature exists alone, but is present in very great quantity, combined with several other elements, in the structures of animals and plants, and combined with oxygen in water; and it is probable that, in the geological ages that preceded that in which living structures were created, all of the hydrogen of the globe was contained in its water and its watery vapour. It is from water that the chemist now obtains it.

Two modes are commonly followed. One of them consists in passing steam over red-hot iron, as shown in *fig.* 21., in which *a* is a retort containing water made to boil by the lamp underneath it. The steam proceeding from this enters the iron tube *b c*, which is filled with iron wire, and kept red-hot by means of the furnace placed around it; *e* is a tube leading from this to the pneumatic trough, and through which the gas flows freely. The rationale of the process is clear. The iron wire is in a metallic state, and consists of pure Fe. The watery vapour consists of $8\,O + 1H$; and as it is passing over the red-hot metal, the O unites with the Fe, and forms oxide of iron, or FeO. The disengaged H then passes along the tube.

Fig. 21.



This mode of decomposing water is, however, a troublesome one, and hydrogen is more conveniently procured by pouring water and some sulphuric acid into a flask into which a number of small bits of zinc

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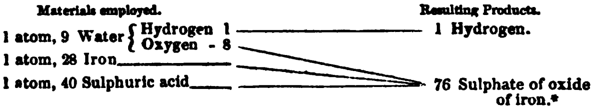
[*] The proper proportion is eight parts of the water to one of acid.

or iron have been introduced. A bent tube perforates the cork of this flask, and leads to the pneumatic trough. A violent effervescence immediately begins, and hydrogen gas is rapidly evolved and collected into a suitable vessel at the pneumatic trough.

The decompositions that take place are as follows: —


Materials employed.
Resulting Products.
1 atom, 9 Water { Hydrogen 1 / Oxygen - 8
1 atom, 28 Iron
1 atom, 40 Sulphuric acid
1 Hydrogen.
76 Sulphate of oxide of iron.*

Hydrogen gas prepared in this manner is never quite pure, as both the iron and zinc of the shops always contain carbon, a little of which comes over with it. But for ordinary purposes it is pure enough.

Hydrogen, when uncombined, is always a gas, no artificial means having yet liquefied it. It is without colour, taste, or smell. It is the lightest of all substances, being nearly fifteen times lighter than air, fifty cubic inches of it only weighing one grain. Hence, of all substances, it is the most suitable for filling balloons, and the first gas-balloons were filled with it. But as the gas which we use to illuminate our towns with is, although a great deal heavier than hydrogen, lighter than atmospheric air, balloons in this country are now made much larger than a hydrogen balloon need be, and filled with it.

Hydrogen cannot support combustion, although, when in the presence of air or oxygen, it is very inflammable. But its relations to flame must be deferred until we have become acquainted with oxygen and carbon.

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* The sulphate of oxide of iron remains in the flask.

Its atomic weight is 1, its symbol H, its density 69·26, and its combining measure two volumes □□.

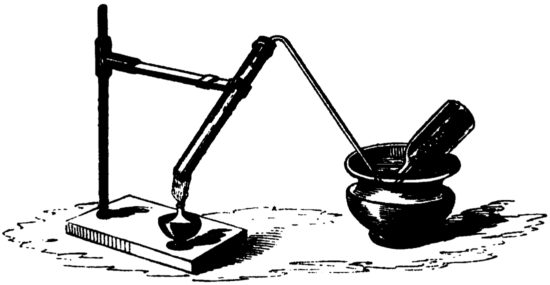
II. Oxygen.

Oxygen was discovered in 1774, by Dr. Priestley, and was by him named dephlogisticated air. A year or two afterwards, Scheele, who was quite ignorant of what Priestley had done, also discovered it, and called it empyreal air. Its present appellation, oxygen, was given to it by Lavoisier, and is derived from *ὀξὺς acid*, and *γεννάω to generate*, because he considered it the sole cause of acidity. We now know, however, that there are acids that do not contain any of it.

Oxygen is the most widely diffused of all the elements. One fifth of the air, eight ninths of water, and about one half of the rocks, and of the structures of animals and plants, consist of it.

Occurring so abundantly as it does, the chemist experiences no difficulty in separating it in a pure state. Usually he disengages it from some oxide or salt containing it, by applying heat. Priestley, when he discovered it, employed oxide of mercury. On applying heat to this substance, it separates into metallic mercury and oxygen, its two constituents. Oxide of mercury, however, is far too expensive for being extensively employed, and another salt, the chlorate of potassa, is the most suitable substance for supplying oxygen. If a little be put at the bottom of a tube, as in *fig.* 22., having a smaller one fitted into it by means of a cork, and a spirit lamp applied

Fig. 22.

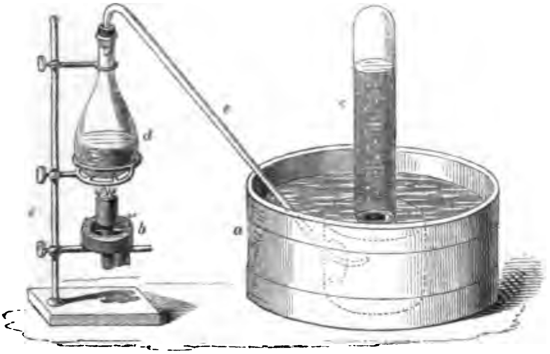


to the tube, the salt quickly melts; it then begins to effervesce, owing to the oxygen it contains separating in bubbles. What remains in the tube is a salt, and the whole of the oxygen passes away through the smaller one into the receiver as a colourless invisible gas. Its presence, however, may be proved by thrusting into the tube a piece of wood with a red-hot top, which will immediately burst into a full flame.

The change that takes place in this process is very intelligible. Besides oxygen, chlorate of potassa $(KO)+(Cl+5O)$ contains chlorine and potassium. When it is heated, the six atoms of oxygen are given off, the chlorine and potassium unite together, forming chloride of potassium, KCl.

When we wish to obtain oxygen in large quantities, we employ a retort or flask in place of a tube, and collect it in jars at the pneumatic trough. The whole apparatus for this purpose is extremely simple, and is figured in *fig.* 23.; *a* is a box or trough*, with a beehive shelf in it, having a hole at the top, and a hole at the side — when it is used, water is poured in until the shelf is covered; — *c* is a jar which has

Fig. 23.



been filled with water and inverted over the hole in the top of the chest; *d* is a flask containing chlorate of potassa; and *e* a tube coming from it through the side hole in the shelf, and terminating just underneath the top hole. When the heat is applied to the flask by the lamp *b*, the separated oxygen passes along the tube, and, being lighter than the water, ascends into the jar, displacing the water. This goes on until all the water is displaced from the jar, which is now full of oxygen. We then slip this jar of oxygen into a little tray containing water, and examine its properties and experiment with it at our leisure.

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* The one figured is Griffin's, made of stoneware.

Oxygen has neither taste, colour, nor smell, and, as it is transparent, it is invisible. It is a little heavier than common air in the proportion of about 11 to 10. Uncombined it is always gaseous, no amount of cold or pressure being able to liquefy it. It unites most abundantly with most of the other simple substances, and can be made to unite with every one of them. These compounds are called oxides or acids. The latter are those compounds of it that have a sour taste, change vegetable blues to red, and vegetable brown to yellow, and form salts by combining with bases. Oxides want all these characters, and many of them are basic, *i. e.* unite with acids to form salts. The act of combining with oxygen is called oxidation, and a body so acted upon is said to be oxidated. Sometimes this oxidation takes place slowly, as when iron combines with the oxygen of the air, or becomes rusty; but upon other occasions it is produced with great rapidity, and attended by the evolution of heat and light. For example, ordinary combustion is nothing else but rapid oxidation.

Oxygen is necessary for the respiration of all animals. Indeed, when we come to Book III. we shall find that respiration is a slow oxidation of the fat of the body. It is necessary, however, that the oxygen for this purpose be diluted as it is in the atmosphere. If pure oxygen be taken into the lungs, the oxidation goes on too rapidly, approaching to combustion, and an animal that so breathes it soon dies.

All substances that are capable of burning in the open air do so with far greater brilliancy in oxygen. A piece of wood with the least spark upon it, introduced into a jar of oxygen, immediately bursts out into a brilliant flame, and phosphorus ignited amid oxygen burns with a flame so brilliant that the eye can scarcely endure to behold it. Even iron wire burns with great brilliancy in this element.

Two results always happen in these cases of combustion, — the oxygen in the jar disappears, and the substance burned increases in weight; it has, in fact, become oxidated. Combustion, it must be remembered, is the same phenomenon, whether occurring in the open air, or in pure oxygen, inasmuch as air is, as we shall afterwards see, essentially diluted oxygen, and combustion in the two differ in degree only, and not in kind. It may seem strange to say, that the heated body by burning increases in weight. And yet it is quite true; and if we took a candle, weighed it, burnt it, and collected the gases that fly off, we should find that they exceed in weight the original weight of the candle. Moreover, if we managed the combustion in a closed jar, or the like, and weighed the quantity of the oxygen that it contained before and after the combustion, we should find that the weight gained by the candle was exactly equivalent to the weight lost by the oxygen.

Although, when a body is rapidly oxidated, heat and light are evolved, it is proper to bear in mind, that the same results are produced during other chemical processes. Thus phosphorus, nay, even copper, will burn in chlorine, potassium take fire in cyanogen, and sulphur and iron filings, when heated to the melting point of the latter, emit intense heat and brilliant light. Nevertheless, in the common meaning of the term, the word combustion means the rapid oxidation of combustible matter, attended by the evolution of heat and flame. And, as we shall afterwards have occasion to notice, the combustible matter is almost invariably one rich in the elementary substance called carbon.

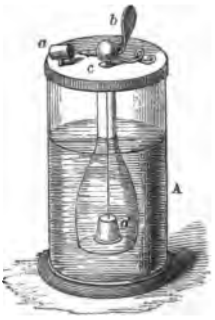
We may here consider the properties of hydrogen, the element before noticed, with regard to flame.

Hydrogen cannot support combustion; and if a lighted candle be put into a jar of it the light is at once extinguished. Itself, however, if in the presence of air (or oxygen), and ignited, burns with a very pale yellow flame of remarkably little illuminating power, but which gives out a most intense heat. The result of the burning the hydrogen in the air (or oxygen) is the formation of water; and for every grain of hydrogen that is burned nine of water are produced. That is to say, every grain of hydrogen that is burned unites with eight grains of oxygen to form water, the result being, of course, nine grains of that compound fluid.

If, however, the hydrogen be not ignited at a jet where it is unmixed with oxygen, but if, before the light be applied, either some of that element or air be mixed with it, instead of burning, the mixture violently and all at once explodes, the explosion being preceded by a sudden flash of light, and attended by a loud noise. The cause of the noise is probably the concussion of the air occasioned by the sudden formation of water, which water, by the great heat that is present, is immediately converted into steam.

It is a curious fact, that if oxygen and hydrogen be mixed together in a jar, they may be preserved for any length of time without combining. The application of flame, as in the above instance, however, immediately produces their union (forming water), and so does the transmission through the mixture of an electric spark. There is another substance that most remarkably facilitates the oxidation of hydrogen, although it is impossible to explain its action. That substance is spongy platinum, the general properties of which will be mentioned hereafter. This power of spongy platinum has been applied to the construction of a philosophical toy,—Dobereiner's hydrogen lamp, *fig*. 24., which illustrates the property in question. In it *a* is a piece of spongy platinum covered with paper; *b* a stop-cock moved at will by the crank *c*; the vessel A, from which this stop-cock proceeds, contains water, sulphuric acid, and zinc *d*, and therefore generates hydrogen. On turning the stop-cock, the hydrogen rushes out, and mixes with the oxygen contained in the air about the platinum; the two immediately combine, and as much heat is evolved as serves to light the piece of paper.

Fig. 24.



Although the flame of burning hydrogen is but little luminous, yet the strongest artificial light that we can produce is obtained by introducing into its flame some infusible substance that it may raise in temperature. Quicklime is generally employed for this purpose, and the hydrogen is kept in full combustion by means of throwing in a jet of oxygen. The heat thereby produced is so great, that the lime becomes so incandescent as to almost rival the sun. At night a ball of lime thus heated is visible for nearly a hundred miles.

A modification of this experiment is used in the oxyhydrogen blowpipe. Before the flame of this, pipeclay, steel, copper, and all the ordinary metals are at once melted, and even the most obdurate of the metals, platinum, yields to its intensity. It is essentially composed of two jets, one leading from a gasometer containing oxygen, and the other from one containing hydrogen, care being taken that the two are immediately united together and do not form a mixture before the flame is applied, in which case an explosion would take place that would probably be fatal to all near.

Its atomic weight is 8, its symbol O, its density 1105, and its combining measure 1 volume □.

COMPOUNDS OF OXYGEN WITH HYDROGEN.

These are two, namely: —

1. *Water.* Properly speaking, water only occurs in the laboratory of the chemist, the water of nature containing, in addition to it, many other substances derived from the inorganic and the organic worlds. The consideration of water is therefore deferred to an after part of this work. One or two of its properties must, however, be noticed here. It is the

standard of density for fluids. Each atom of it is composed of one of oxygen and one of hydrogen, and it unites with many other substances; sometimes in almost any proportion, as in cases of common solution, but at other times in fixed definite proportions, according to its atomic constitution. When it forms this latter class of compounds, the term hydrate is applied.

Its atomic weight is 9, its symbol HO (or often Aq, from *aqua*), its density 1000.

2. *Peroxide or binoxide of hydrogen.* This compound was discovered by Thenard in 1818, and could a cheap supply of it be obtained, would, doubtless, be of great service in the arts. Its preparation, however, is very difficult and costly, and, moreover, it is decomposed when exposed to a temperature of more than 59°. It possesses very great bleaching properties.

Its atomic weight is 17, and its symbol $H+2O$.

* * *

III. CHLORINE.

This elementary substance was discovered by Scheele in 1774, and was by him considered a compound body, and named dephlogisticated marine acid. The French chemists subsequently named it oxymuriatic acid, from an idea that it was a compound of muriatic acid and oxygen. Gay Lussac and Thenard suspected its elementary nature; but the discovery of such was reserved for Davy, and it was he that gave it the name it now bears, chlorine, from χλωρὸς, *green*, owing to the green colour of its vapour.

Chlorine exists abundantly in nature, generally combined with sodium, forming chloride of sodium or common salt. This compound is present in immense quantities in the waters of the ocean, forms large mineral deposits, and is to be found in every soil and in the structure of almost every animal and plant. Chlorine likewise exists, but in less quantities, in other combinations.

On a small scale and for examination in the laboratory, chlorine is most conveniently obtained by decomposing hydrochloric acid by means of peroxide of manganese. If these substances are introduced into a retort, chlorine is immediately evolved, and in still greater quantities if heat be applied to the retort.

The nature of the change that takes place is very simple. The composition of the two substances and of the products obtained from them is as follows:— The composition of the substances is first put down.

Manganese -	27	or 1 atom	Mn.	Chlorine -	70	or 2 atoms	
Oxygen -	16	or 2 atoms	2O	Hydrogen	2	or 2 atoms	H.
Perox. of manganese	43	or 1 atom	Mn 2O.	Hydrochlor. acid	72	or 2 atoms	2(Cl H.)

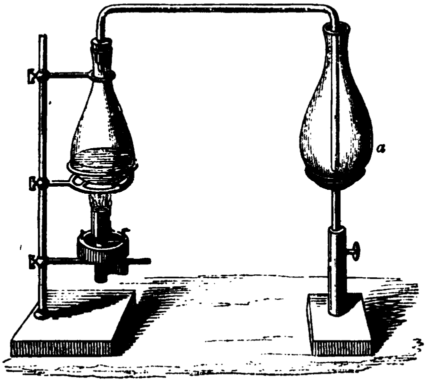
The results are—

Manganese - -	27	Hydrogen	2		Chlorine, 35. 1 atom.
Chlorine - -	35	Oxygen	16		
Chloride of manganese	63	Water -	18	2 atoms.	

The chloride of manganese and the water remain in the retort, and the one atom of chlorine that is liberated passes in a gaseous form into the receiving jar. It happens that cold water absorbs about twice its weight of chlorine, and hence that fluid cannot be used in the pneumatic trough. Generally speaking, when water absorbs a gas, the chemist collects that gas over mercury; but for collecting chlorine mercury is more objectionable than water, inasmuch as chlorine and mercury combine together and form calomel. Warm water, however, does not absorb chlorine, and, therefore, this element is collected in bottles containing water heated to about 90°.

There is no necessity, however, for using a pneumatic trough at all in collecting chlorine. If, as in *fig.* 25., the tube from the flask containing the acid and the manganese be brought to the bottom of a dry flask as *a*, the chlorine being much heavier than the air pushes out the air, just as water poured in would do. The colour of the gas allows us to see quite well when the flask is full. The flask's mouth should then be closed with a well-ground glass stopper.

Fig. 25.



Chlorine has a greenish-yellow colour and a peculiar marine smell. Breathed in any quantity it is one of the most suffocating of the gases, exciting violent spasm and irritation of the glottis. Diluted, it has no such effect, and consumptive patients are sometimes ordered to inhale it. Under the pressure of four atmospheres it becomes fluid. This fluid chlorine is of a bright yellow colour, which does not freeze on being exposed to the cold of zero, and which rapidly becomes a gas again whenever the pressure is taken away.

As mentioned, cold water absorbs twice its volume of it, and a solution made by passing chlorine through cold water possesses the colour, taste and most of the properties of the gas itself. This solution, when heated, parts with its chlorine.

Light has no action upon dry chlorine; but if water be present the chlorine decomposes that fluid, and unites with its hydrogen, forming hydrochloric acid, and liberating oxygen. This decomposition takes place most rapidly in bright sunshine, and not at all in the dark. Hence it is necessary to keep moist chlorine gas, or its aqueous solution, in the dark.

Chlorine is a supporter of combustion. Strange to say, however, the substances, *i. e.* those rich in carbon, that burn so vigorously in oxygen (the supporter of combustion proper), scarcely burn at all in chlorine. If a lighted candle, for instance, be plunged into chlorine, it burns for a very little with a pale blue flame and then goes out. In this case none of the carbon is consumed, but is all deposited as charcoal. Phosphorus takes fire in it spontaneously, and so do many of the metals, as copper, tin, &c., if introduced into the chlorine in the state of powder or fine leaves. In all these cases the combustible substances unite with chlorine.

The affinity of chlorine for hydrogen is very great; and the two gases, if mixed together and exposed to the light, unite. As just stated, chlorine in this manner decomposes water, and indeed it decomposes almost every organic substance containing hydrogen

From this last-mentioned property of chlorine, it is the best substance that we possess for removing noxious effluvia and putrid odours. It decomposes the water that these contain, unites with its hydrogen, and the oxygen that is set free oxidises the organic matter. In this manner the composition, and therefore the peculiarities of, the noxious effluvias are changed. It has been supposed that it can also decompose contagious miasmata, but this is by no means certain.

But the most valuable property that chlorine possesses is its power of thoroughly removing all animal and vegetable colours. But the consideration of this action is deferred until Book III., where the colouring matters derived from the organic world will be noticed.

The compounds of chlorine that are not acid are termed chlorides.

Chlorine may be readily distinguished from all other substances by its bleaching property, and by the fact that a solution of nitrate of silver occasions in it a dense white precipitate, that becomes dark on exposure to the air, is insoluble in acids, but dissolves in ammonia. This substance is the chloride of silver, hereafter to be noticed.

Its atomic weight is about 35, its density 2440, its symbol Cl, and its combining measure two volumes ☐☐.

The composition of the compounds of chlorine described in this section are,—

					Atomic Weight.	Formula
1. *Hydrochloric acid*	—Chlorine	35	—1 atom	+Hydrogen 1, 1 atom	=36	H+Cl.
2. *Hypochlorous acid*	do.	do.	do.	+Oxygen - 8, 1 atom	=43	Cl+O.
3. *Chlorous acid* -	do.	do.	do.	+Oxygen - 32, 4 atoms	=67	Cl+4O.
4. *Chloric acid* -	do.	do.	do.	+Oxygen - 40, 5 atoms	=75	Cl+5O.
5. *Perchloric acid* -	do.	do.	do.	+Oxygen - 56, 7 atoms	=91	Cl+7O.

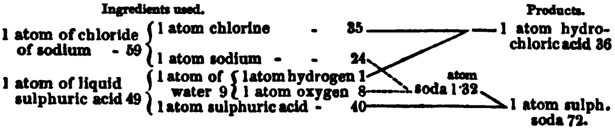
1. Hydrochloric Acid.

This compound naturally exists in a gaseous state, but its solution in water has been known to the alchemists and chemists for at least a thousand years. It has been at various times named spirit of salt, marine acid, and muriatic acid. Since its true nature has been known it has borne the name of hydrochloric acid; but the best name for it would be chloride of hydrogen.

As mentioned above, this compound is formed when hydrogen and chlorine are mixed together, and exposed to solar light, and also whenever chlorine comes in contact under the same circumstance with organic matter containing hydrogen.

In the laboratory hydrochloric acid gas is obtained by the action of strong liquid sulphuric acid upon dried chloride of sodium. These ingredients are put into a retort, and the gas collected over mercury; or, as this gas is heavier than air, it may be conveyed along a tube after the fashion of chlorine mentioned at p. 107.

The decomposition that takes place is as follows:

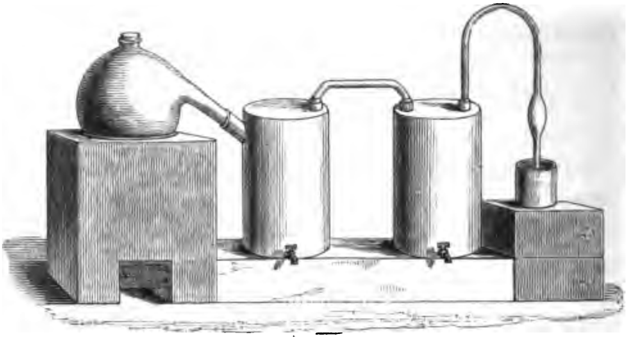


Hydrochloric acid thus obtained is a colourless, invisible gas, fuming in the air in consequence of its affinity for water, a substance that very rapidly absorbs it. It may be liquified by great pressure (40 atmospheres). When it comes in contact with an oxide of a metal, as the oxide of copper, for instance, both are decomposed, the hydrogen and oxygen unite to form water, and the chloride of copper is produced. Its fumes are very destructive to both animal and vegetable life. It is of little or no use.

Liquid muriatic acid is prepared in the laboratory by mixing chloride of sodium, sulphuric acid, and water, and then distilling, the *rationale* of the action that takes place being very evident; or the commercial acid is purified.

The common mode of making muriatic acid in commerce is to employ an iron or a stoneware pot set over a fireplace, having a stone head luted to it, which is connected with double-necked stone bottles, the last having a safety tube dipping into water; see *fig.* 26. The ingredients employed are the same as those above mentioned. Carbonate of soda is now extensively made from sulphate of soda, the sulphate being formed by this process. Sometimes the hydrochloric acid produced in this operation is preserved, but more generally it is allowed to run to waste, the demand for it being so small.

Fig. 26.



Liquid hydrochloric acid varies in specific quantity according to its strength. The following table, taken from Thomson, exhibits the relative quantities of gas and water in specimens of different specific gravity:—

Specific Gravity.	Real Acid in 100 of the Liquid.	Atoms of Acid.	Atoms of Water.
1·203	40	1	6
1·179	37	1	7
1·162	34	1	8
1·149	31	1	9
1·139	29	1	10
1·119	25	1	12
1·100	21	1	15
1·078	17	1	20

A strong solution of hydrochloric acid is, when pure (which the commercial acid never is), colourless; when exposed to the air it emits white vapours; it is intensely sour, reddens litmus paper strongly, and neutralises alkalies. It freezes at −60°, and if heated up to 110°, boils, that is, the gaseous acid separates. Some of the metals decompose it, hydrogen gas being evolved, and a chloride being formed in solution. Several oxides and acids also decompose it, water and a chloride being produced.

Chemically, hydrochloric acid is distinguished by forming with nitrate of silver the same substance that chlorine does, while we know that it is not chlorine from the absence of bleaching properties.

To the chemist, hydrochloric acid is a most important substance, but it is little employed in the arts.

The atomic weight of hydrochloric acid is about 36, and its symbol $H+Cl$.

2. HYPOCHLOROUS ACID.

This substance is prepared by putting into a bottle containing chlorine and water some powdered red oxide of mercury, corking the bottle and shaking it well. A portion of the chlorine combines with the mercury, and forms calomel, *i. e.* chloride of mercury, and the remainder unites with the oxygen, forming the acid under notice, and which dissolves in the water.

This solution of it is yellow, with a strong chlorine colour, and it possesses very powerful bleaching properties. It combines with bases to form salts which possess the bleaching qualities of the acid to a very great extent, and some of which, when we have considered organic colours, will again occupy our attention.

3. Chlorous Acid.

This is now often called hypochloric acid. It is procured by the action of sulphuric acid upon chlorate of potassa, a compound alluded to when considering the preparation of oxygen. It is a very explosive substance, and forms salts which possess bleaching properties; but as neither it nor its salts are ever put to any practical use, we may be content to merely notice it and them.

4. Chloric Acid.

Chloric acid is obtained by adding sulphuric acid to a solution of chlorate of baryta, upon which the insoluble sulphate of baryta falls to the bottom, leaving the chloric acid free.

It possesses no bleaching properties, and is in itself of no consequence; but it forms with potassa a salt, the chlorate of potassa, which ignites when rubbed against a hard substance with sulphur, and which constitutes the illuminating end of lucifer matches (*see* chlorate of potassa afterwards).

5. Perchloric Acid.

This may be obtained by decomposing perchlorate of potassa; but as it is of no practical use, we may pass it over. It is distinguished by forming with potassa a salt little soluble in water.

IV. BROMINE.

This elementary body was discovered in 1826 by Balard of Montpellier. In its chemical properties it is very analogous to chlorine, the element mentioned above, and to iodine, the element that will next occupy our attention. It, too, in nature, is almost always found associated with one or both of them; and this would seem to have been also the case in far earlier geological ages than our own. As yet, neither bromine nor its compounds have been put to practical uses; but it is probable that they may be.

The name is derived from *βρῶμος, a rank smell,* on account of its unpleasant odour. It exists, but in very small quantities, in sea water, in the form of either bromide of sodium or magnesium, particularly the latter; in some mineral springs, and in some marine animals and plants.

In this country bromine (which, however, is prepared in very small quantities, and is very dear) is extracted from sea water. The old-fashioned plan of obtaining common salt was to evaporate sea water, removing the crystals of salt as they were deposited. The mother water that was left was a dense oily liquid (containing all the marine salts except the chloride of sodium), called bittern; and from this bittern bromine can be obtained.

To do so, chlorine is sent through the bittern, which separates the bromine from the magnesium, and sets it free among the bittern, which it immediately tinges yellow. The liquid is then put into a retort and heat is applied, upon which the bromine distils over into the receiver.

Bromine is the only element, with the exception of mercury, that is fluid at ordinary temperatures. Its colour is blackish red, its odour unpleasant, and its taste powerful. It is a poison. Its specific gravity is about 3°. At a little below zero it freezes, and then forms a brittle solid, and it boils at 116°. It is soluble in water, alcohol, and ether. It bleaches readily. A lighted candle introduced into its vapour, burns for a little, with a pale blue flame, and then goes out. It corrodes the skin. Mixed with starch it forms a yellow compound.

Its atomic weight is 78, and its symbol Br.; its combining measure □□.

The compounds of bromine described, or rather enumerated, in this section are the following: —

		Formula.
1. *Hydrobromic acid*	Bromine 1 atom 78 + Hydrogen 1 atom 1 = 79	$H + Br$.
2. *Bromic acid* -	Bromine 1 atom 78 + Oxygen - 5 atoms 40 = 118	$Br + 5O$.

1. Hydrobromic Acid.

If bromine vapour and hydrogen be mixed together and exposed to even strong solar rays, they do not unite. On introducing a piece of hot iron, however, the particles in the immediate vicinity to it do, but the change does not extend to the whole mixture.

Hydrobromic acid strongly resembles, in its properties and relations, hydrochloric, from which, however, it is easily distinguished by evolving bromine on the addition of either chlorine or nitric acid. It is a gas, which is abundantly absorbed by water: it does *not* decompose water. It acts, however, upon metallic oxides in a manner exactly analogous to hydrochloric acid.

2. BROMIC ACID.

This compound has been little examined, is of small consequence, and may be passed over.

* * *

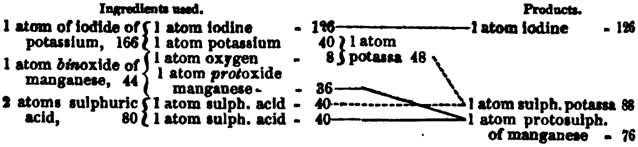
V. IODINE.

Iodine was discovered in 1811, by M. Courtois, a saltpetre manufacturer at Paris; but its real nature was determined by Davy and Gay Lussac, who proved that it was a non-metallic element, very analogous in its chemical relations to chlorine. It was named iodine, from ἰωδὴς, *violet-coloured*, on account of the colour of its vapour.

Iodine is found, but very sparingly indeed, in some metallic ores; many mineral waters contain it; but its main source is the waters of the ocean, where it is combined with sodium or potassium. It is an essential constituent of many marine plants, particularly of plants belonging to the order *Algæ* and it is a compound part of the structure of some marine animals that feed upon these sea-weeds. The common cod is an example of these latter. In this animal the iodine is mainly or entirely confined to the liver.

For commercial purposes iodine is always extracted from *kelp*, *i. e.*, the ash of sea-weed. This kelp is a very heterogeneous mass, and contains chlorides of sodium and potassium, sulphates of soda and potassa, and iodides of sodium and potassium. The first part of the process consists in getting rid of the other salts, which are in much greater quantities than the iodides. This is done by repeated crystallisation, as the iodide, being very soluble, is retained in the mother liquid. This mother liquid at length becomes a dark, thick mass, and is called the iodine ley.

This is put into a stoneware still, sulphuric acid and oxide of manganese are added to it, and heat is applied. Iodine then distils over, and is then washed with water, and dried between pieces of bibulous paper.



In this process 2 atoms, or 80, of sulphuric acid, react on 1 atom, or 44, of binoxide of manganese, and also on 1 atom, or 166, of iodide of potassium; and yield 1 atom, or 126, of iodine, 1 atom, or 88, of sulphate of potash, and 1 atom, or 76, of sulphate of *prot*oxide of manganese.

Iodine, at common temperatures, is a soft, friable, and opaque solid, of a bluish-black colour, and has a semi-metallic lustre. It is usually met with in micaceous scales; but it sometimes crystallises in large rhomboidal plates. Its taste is hot and acrid, and its odour heavy, disagreeable, and peculiar. It freezes at 225°, and volatilises at 347°, its vapour having a peculiar and characteristic violet hue.

It is very sparingly soluble in water, requiring about 7000 times its weight of that fluid to dissolve it. Even in this quantity it communicates a brown tinge to the water. It dissolves readily in alcohol.

Besides its peculiar and violet-coloured vapour, iodine is characterised by forming, with starch, a deep blue compound, insoluble in cold water. This test is so delicate, that a liquid containing so small a quantity as $\frac{1}{450,000}$th of its weight of iodine strikes the blue colour when a little starch is added to it. It is necessary that the iodine be in a free state, and the solution cold, as hot water dissolves the blue compound, and forms a colourless solution.

The atomic weight of iodine is 126, its symbol I, and the combining measure of its vapour □□.

The compounds of iodine described in this section are —

		Formula.
1. *Hydriodic acid* -	Iodine 1 atom 126+Hydrogen 1 atom 1=127	$H+I$.
2. *Iodic acid* - -	Iodine 1 atom 126+Oxygen - 5 atoms 40=166	$I+5O$.
3. *Chloride of iodine*	Iodine 1 atom 126+Chlorine - 1 atom 35=161	$I+Cl$.
4. *Bromide of iodine*	Composition uncertain.	

1. HYDRIODIC ACID.

This is the most important compound of iodine. It may be formed by the direct union of its elements, when a mixture of hydrogen gas and iodine vapour are transmitted through a red-hot porcelain tube. It is a gas having very acid properties, and forms salts with alkalis called hydriodates. It is rapidly absorbed by water, and therefore cannot be collected over that fluid.

2. Iodic Acid.

One of the modes of obtaining this acid is, by digesting iodine in nitric acid. In this case the nitric acid parts with a portion of its oxygen, and becomes nitrous acid, while the liberated oxygen combines with the iodine. It is a solid, semi-transparent, soluble, deliquescent, acid substance. Its salts are called iodates. With charcoal, sugar, and other combustibles it forms compounds that detonate on being heated.

3. Chloride of Iodine.

There is more than one chloride of iodine; but they are only interesting as being employed in one mode of procuring iodic acid. It is not necessary to attend particularly to them.

4. Bromide of Iodine.

Bromine and iodine combine together by simply mixing them; but the compound is of little interest.

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VI. Fluorine.

Fluorine has scarcely been obtained in a separate form, but its existence is inferred partly from analogical reasoning and partly from some experiments of the Messrs. Knox. The reason that it cannot be prepared in an isolated state is, that its affinity for other substances is so great, that no sooner is it
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separated from its combination with one substance than it enters into union with some other, attacking the materials of which the apparatus used is composed.

This element exists pretty extensively in nature in combination with calcium, forming fluoride of calcium, or fluor spar. Its atomic weight is calculated to be about 18; its symbol is F.

The only compound of fluorine necessary to notice here is—

		Formula.
Hydrofluoric acid -	Fluorine 1 atom 18+Hydrogen 1 atom = 19	H+F.

This acid is obtained by mixing together fluor-spar and sulphuric acid, and by applying heat to a leaden retort in which the mixture is placed. Sulphate of lime and hydrofluoric acid are formed, the latter of which passes off in a gaseous form into the receiver. This is kept cold by means of ice, and the acid becomes fluid. The chemical changes are precisely the same as in the formation of hydrochloric acid, fluorine being substituted for chlorine.

At 32° this acid is fluid, and remains in that state, if kept in well-closed bottles, up to 60°. It is heavier than water, and its density may be still further increased by adding a little water to it. Its affinity for this fluid is very strong, and its combination with it is attended by a loud hissing noise.

Of all known substances hydrofluoric acid is the most destructive to animal textures. A small drop, the size of a pin's head, instantly destroys the skin, and induces a deep ulcer. Hence, in its concentrated form, it is a very dangerous thing to have anything to do with. It acts very energetically upon glass, and is used in etching, as will be considered in a subsequent page. For this reason it cannot be prepared or kept in bottles of glass; and lead, platinum, or silver are employed for the purpose.

Hydrofluoric acid has a strong sour taste, reddens litmus paper, and neutralises alkalis, sometimes forming with them hydrofluates, but more commonly producing a metallic fluoride and water.

When brought into contact with certain metals it is decomposed, hydrogen gas being given off, and a metallic fluoride formed. Upon potassium it acts in this way with great energy, and an explosion attended with heat and light ensues.

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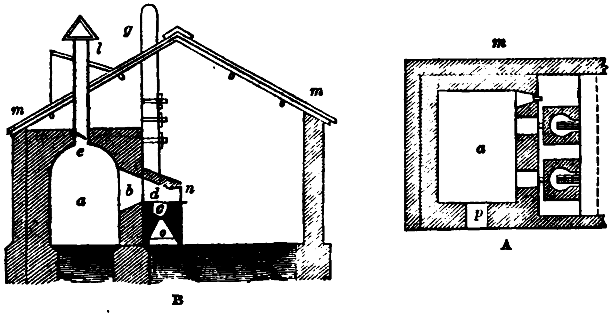
VII. Sulphur.

This substance has been known from the remotest antiquity, and is mentioned by Moses, Homer, and other antient writers. It occurs in nature abundantly, both native and in a state of combination with other elementary bodies. Native, or virgin sulphur as it is sometimes called, is found in the neighbourhood of volcanos; and our principal commercial source of it is Sicily. Such is probably derived from the decomposition of metallic sulphurets from the high temperature in the interior of the earth. The most important ores of copper, lead, and many other metals are sulphurets; and a great portion of the sulphur of commerce is obtained from the sublimation of iron pyrites, or the bisulphuret of iron. Sulphur also exists in
~~~~~~~~

combination with metallic oxides and oxygen, forming sulphates; and it is likewise a necessary constituent of most animals and plants. The common name of sulphur is brimstone.

Native sulphur is purified by melting it; and both it and that obtained by the sublimation of iron pyrites are, when molten, poured into moulds, in which it solidifies. It then constitutes roll sulphur. Flowers of sulphur, as they are called, are prepared by condensing sulphur vapour so rapidly that the atoms have not time to form crystals of any perceptible size, but appear like a soft powder, although this is in fact composed of a number of minute crystals. To do this the apparatus here figured (*fig.* 27.) is employed. A and

Fig. 27.



B are two sections; the former a vertical, the latter a horizontal one. The same letters refer to both. In an apartment and shed *m m*, is a vaulted shed *a*, having a capacity of at least two thousand cubic feet. Outside of *a* is *c*, an iron pan, filled with sulphur kept boiling by the fire at *o*. Both boiler and fireplace are lined with brick, to prevent much heat being communicated to the chamber *a*. The draught from the fire passes to the chimney *g*. The pan is supplied with sulphur at the door *n*, and the vapour of the sulphur mixes with the air in the wide space *d*, and, passing through *b*, rises into the shed or chamber, where it meets with a quantity of cold air and immediately falls like a snow shower on the floor. This powder is removed by the door at *p*. If the mixture of sulphur vapour and air should inflame, the valve *e* is opened by the explosure, and the gases escape at *l*.

Sulphur is prepared in a state of even more minute division. This preparation is called *lac sulphuris*, or milk of sulphur, and was known to Geber. It is prepared by decomposing sulphuret of potassium, or calcium, by means of hydrochloric acid. When this is done the sulphur falls down as a purely white powder. It generally, perhaps essentially, contains a little water.

Common sulphur is a brittle solid, of a yellowish green colour, and of scarcely any taste. If rubbed, it emits a peculiar odour. At 232° it melts, but if heated up to 320° it begins to thicken, and to acquire a reddish tint; and at temperatures between 429° and 482° it becomes almost solid. From 482° to the temperature at which it begins to volatilise, it again becomes fluid.

Another remarkable physical property of sulphur is, that if heated to 430°, and then suddenly cooled, by being thrown into water, it becomes a ductile mass, and may be used for taking the impressions of seals. After a time it returns to its common state.

At 550° sulphur volatilises freely. During its volatilisation a little sulphurous acid is always formed. In practice, this impurity is got rid of by washing the sulphur, for sulphur itself is altogether insoluble in water.

When sulphur is heated in the open air, it takes fire at about 300°, burning with a blue flame. The sulphur combines with the oxygen of the air, forming sulphurous acid.

Sulphur is extensively employed. It is used in medicine; is the substance almost univerally adopted for making all kinds of matches ignite more readily; in making casts; and in the manufacture of gunpowder, of sulphurous and of sulphuric acids.

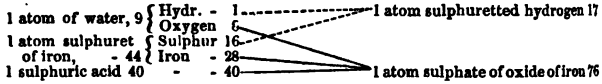
The atomic weight of sulphur is about 16, its specific gravity 1·9, its symbol S, and its combining measure □.

The compounds of sulphur that require mention here are —

1. *Hydrosulphuric acid*	-	Sulphur	1 atom	16+	Hydrogen	1 atom	1=17	$H\,S$.
2. *Persulphuret of hydrogen*		Sulphur	2 atoms	32+	Hydrogen	1 atom	1=33	$H+2S$.
3. *Sulphurous acid*	-	Sulphur	1 atom	16+	Oxygen -	2 atoms	16=32	$S+2O$.
4. *Sulphuric acid* -	-	Sulphur	1 atom	16+	Oxygen -	3 atoms	24=40	$S+3O$.
5. *Hyposulphurous acid*	-	Sulphur	2 atoms	32+	Oxygen -	2 atoms	16=48	$2S+2O$.
6. *Hyposulphuric acid*	-	Sulphur	2 atoms	32+	Oxygen -	5 atoms	40=72	$2S+5O$.
7. *Chloride of sulphur*	-	Sulphur	2 atoms	32+	Chlorine	1 atom	35=67	$2S+Cl$.
8. *Bromide of sulphur*	-	Composition uncertain.						
9. *Iodide of sulphur*	-	do.						

1. HYDROSULPHURIC ACID.

The common name of this compound is sulphuretted hydrogen. To prepare it, sulphuret of iron is acted upon by diluted sulphuric acid. An effervescence takes place, sulphuretted hydrogen is evolved, and sulphate of oxide of iron is left in the retort. In this process, the water associated with the acid is decomposed, its oxygen uniting with the iron to form the oxide of iron that unites with the sulphuric acid, and the hydrogen forming, with sulphur, the compound now under notice.


1 atom of water, 9
Hydr. - 1
Oxygen 8
1 atom sulphuret of iron, - 44
Sulphur 16
Iron - 28
1 sulphuric acid 40 - - 40
1 atom sulphuretted hydrogen 17
1 atom sulphate of oxide of iron 76

Sulphuretted hydrogen is a colourless gas, which reddens litmus paper slightly. It is known by its extremely offensive smell, similar to that of rotten eggs, which, indeed, owe their odour to it. Under a pressure of seventeen atmospheres it becomes liquid. It is remarkably destructive to animal life, and the presence of $\frac{1}{1500}$th of it in air has been known to destroy a bird. It extinguishes all burning bodies placed in it, but itself takes fire, and is resolved into water, sulphurous acid, and sulphur.

It is decomposed by chlorine; and hence that substance acts as a disinfectant and purifier of sewers, &c., the offensive nature of which is attributable to the presence of sulphuretted hydrogen, from the decomposition of animal matter. Many of the metals decompose it, particularly when heated; and as the air always contains a certain amount of it, silver and gold utensils gradually become tarnished, that is, they combine with the sulphur of the sulphuretted hydrogen, and hence become coated with sulphuret of gold or silver.

Sulphuretted hydrogen dissolves readily in water, and this solution of it is much used by the chemist as a test for metals, from the solution of which it precipitates sulphurets that have various colours. Thus, sulphuretted hydrogen precipitates—

From a solution of antimony,	*orange,*	sulphuret of that metal.
From a solution of manganese,	*flesh red,*	sulphuret of manganese.
From a solution of arsenic,	*canary yellow,*	sulphuret of arsenic.
From a solution of lead, mercury, &c.,	*black,*	sulphuret of these metals.

Sulphuretted hydrogen is a common ingredient in mineral springs, and in such cases is probably evolved by the action of water upon sulphuret of iron at a high temperature. These are easily known, by their offensive odour, their blackening a silver spoon, or giving, with a soluble salt of lead, the black precipitate of sulphuret of lead.

2. PERSULPHURET OF HYDROGEN.

This compound is believed to consist of two atoms of sulphur and one of hydrogen. It is a vivid yellow fluid, having a density of about 1·7, and of the same odour, although feebler, as sulphuretted hydrogen. The elements composing it are so feebly united, that it always tends to resolve itself into sulphur and sulphuretted hydrogen. The contact, too, of most substances, tends to promote this decomposition. Like the binoxide of hydrogen, to which it is very analogous, the presence of acids renders it much more stable. Like it, too, it whitens the skin, and possesses bleaching properties.

3. SULPHUROUS ACID.

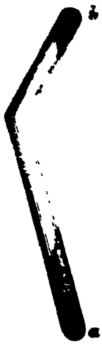
This compound has been known and employed therapeutically from a very remote antiquity. Homer, for example, mentions sulphur fumigations. One of its old names was *spirit of sulphur by the bell.* We are indebted to Dr. Priestley for a knowledge of its composition.

At ordinary temperatures and pressure sulphurous acid is a gas. It is produced whenever sulphur is burned in air or oxygen; or, if required very pure, it is obtained by deoxidising sulphuric acid by means of mercury. In this latter case the metal combines with one atom of the oxygen of the sulphuric acid, upon which sulphurous acid is evolved. As sulphurous acid is very readily absorbed by water, it must be collected over mercury.

Sulphurous acid has a peculiar odour (the smell of burning sulphur), and cannot be breathed. It is colourless and transparent, does not support combustion, and is not itself combustible; and it bleaches a variety of animal and vegetable colours, and hence is used in the arts. The mode in which it bleaches is peculiar. It combines with the colouring matter, forming a white compound; but as this compound is gradually decomposed, and the acid cast off into the air, bleaching with sulphurous acid is not permanent. This compound may at once be decomposed by the addition to it of a stronger acid. Thus, if a red rose be exposed to the fumes of burning sulphur, the sulphurous acid combines with the colour, and the rose becomes white. If this white rose be then washed with a little diluted sulphuric acid, the white compound is decomposed, and the rose becomes red again.

The specific gravity of sulphurous acid is 2·2. Water takes up thirty-three times its volume of this gas. Sulphurous acid may be obtained in a liquid form by putting into a tube (as *fig.* 28.) some sulphuric acid and mercury, and then sealing up the ends. By applying heat to *a*, and cold to *b* (by means of ether), the gas is evolved and liquefied by its own pressure at *b*.

Fig. 28.



Sulphurous acid forms salts with bases, called sulphites. It is a very feeble acid, and is expelled from all its combinations by almost every other acid save the carbonic.

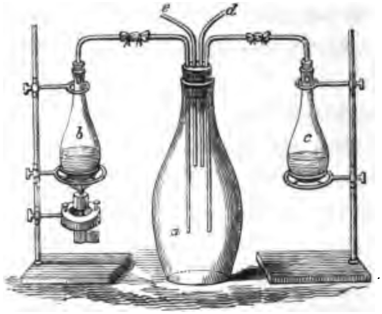
Besides being used in bleaching, sulphurous acid is employed as a disinfectant and in the treatment of skin diseases.

4. SULPHURIC ACID.

This acid was known to Geber as early as the seventh century. In English commerce it almost invariably occurs combined with water, and is a hydrate of sulphuric acid. In common language this is called oil of vitriol. It is, from the energy of its action and the many numerous compounds that it forms, one of the most important of chemical substances.

The preparation of sulphuric acid may be illustrated on a small scale in the following experiment: — *a* (*fig.* 29.) is a central vessel containing common air and slightly moistened with water; *b* is a flask giving off sulphurous acid; and *c* another, giving off nitric oxide into it. By and by the interior of the central vessel becomes covered with a deposit somewhat resembling hoar frost. In order that this may continue to be deposited, fresh air must be introduced into the central flask by blowing into *d*; upon doing which the gases will escape by *e*.

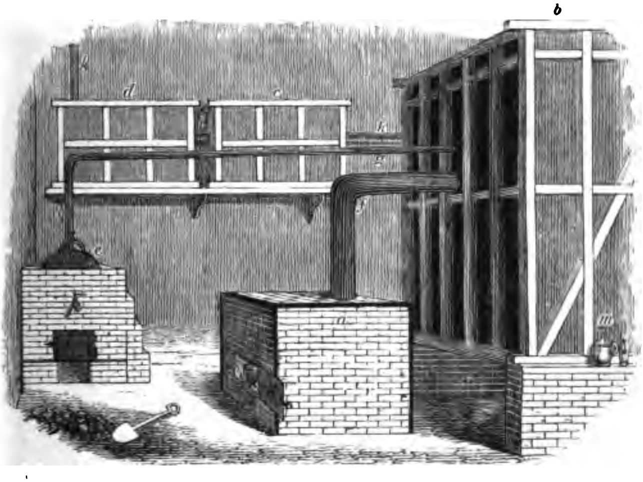
Fig. 29.



This white crystalline substance is a compound of sulphurous acid and nitrous acid; and an atom of oxygen leaves the latter and joins the former, the result being the formation of sulphuric acid and hyponitric acid.

With regard to the manufacture of sulphuric acid on a large scale in this country we extract from Pereira: — "Crude sulphur is burned on an iron plate in a furnace, and the gas (sulphurous acid) thereby produced is conveyed by a pipe or chimney into a leaden chamber, the bottom of which is covered to the extent of four or five inches with water. The furnace door has two apertures to regulate the supply of atmospherical air. After combustion has continued for some hours, an iron pot containing a mixture of the nitrate of potash or soda and oil of vitriol is introduced into the furnace, and steam is conveyed into the leaden chamber; so that sulphurous acid, nitric or nitrous acid (generated by the action of the oil of vitriol on the nitrate of potash and volatilised by the heat produced by the combustion of the sulphur), and steam are conveyed into the leaden chamber, and by their mutual reaction produce sulphuric acid, which is absorbed by the water at the bottom of the chamber. To prevent loss in the process, the residual gases are conveyed through two other small leaden chambers (also containing water) before they are allowed to escape into the air.

Fig. 30.



a, Furnace. *b*, *c*, *d*, First, second, and third leaden chambers. *e*, Boiler. *f*, Pipe of do. *g*, Steam-hole. *h*, *i*, Pipes from first, second, and third chambers. *k*, Waste-pipe. *l*, Man-hole.

"When the liquid in the leaden chamber has acquired a specific gravity of about 1·5, it is conveyed by leaden pipes into rectangular leaden boilers, where it is evaporated and concentrated until its specific gravity is 1·7; but in some manufactories this part of the process is omitted. The final concentration is effected by boiling it down in glass or platinum retorts; the latter, notwithstanding their great price, are now usually adopted. Their price of course depends upon their size. Mr. V. Parker had one which held thirty gallons, and cost about 360*l.*; but sometimes they are made so large that they are worth 1000*l.* each. In this apparatus the acid is deprived of a part of its water and some sulphurous acid; and when it has attained a sufficient degree of concentration it is drawn off by means of a platinum syphon into carboys.

"The theory of the process is the following:—When sulphurous and nitric acids come into contact, the first loses, the second gains, oxygen, by which sulphuric and nitrous acids are formed. Dry nitrous acid and sulphurous acid gases undergo no change when mixed; but if a little moisture be admitted, re-action takes place, and a crystalline compound of sulphuric acid, hyponitrous acid, and water is formed. This, falling into water, is immediately decomposed, sulphuric acid is dissolved, and the hyponitrous acid is resolved into nitrous acid and binoxide of nitrogen, which escape with effervescence in a gaseous form. The nitrous acid thus disengaged, with that produced by the union of the binoxide with oxygen of the atmosphere, becoming mixed with more sulphurous acid and water, give rise to a fresh portion of the crystalline compound. One part of the nitrous acid is converted by the water present into nitric acid and binoxide of nitrogen, while another portion loses the whole of its oxygen, by which a little nitrogen gas is disengaged."[*]

At Nordhausen in Germany another method of obtaining sulphuric acid is followed. It consists in decomposing sulphate of oxide of iron (in common language, green vitriol) by heat. This salt in its ordinary state contains six atoms of water, but when strongly dried by the fire, five of these are driven off and only one remains. If this dried sulphate be then exposed to a red heat, the sulphuric acid passes off along with the water, leaving oxide of iron behind. This hydrate of sulphuric acid is popularly called *fuming sulphuric acid*, and it consists of two atoms of acid united with one of water.

If this last-mentioned acid be introduced into a retort, the receiver of which is surrounded with snow, and heat gently applied to it, a transparent colourless vapour passes over, which condenses into a white crystalline solid. This is anhydrous or pure sulphuric acid.

Anhydrous sulphuric acid is tough and elastic, it liquifies at 66°, and boils at from 104° to 122°. The specific gravity of it at 78° is nearly 2.

The Nordhausen acid is a dark brown oily-looking liquid, that gives out dense fumes when exposed to the air; from which circumstance it derives its common name. As just stated, every two atoms of it are associated with one of water.

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[*] The complicated changes that take place in this process will be unintelligible unless the reader is acquainted with the compounds of oxygen and nitrogen, not described in this work until p. 145, and following.

The oil of vitriol, or English commercial sulphuric acid, is a transparent, colourless, inodorous, very acrid and corrosive fluid. It is a very intense acid, and saturates bases and displaces other acids from their combinations with great power. It has a very strong affinity for water, and the union of the two is attended with evolution of heat and considerable condensation. If exposed to air containing watery vapour, it rapidly absorbs it. Indeed, it is to its great affinity for water that it owes its corrosive property, and applied to animal or vegetable structures it destroys them by separating from them the water necessary to their being.

English sulphuric acid has a specific gravity of about 1·8, and such acid is composed of one atom of water and one of acid; it freezes at $-15°$, and boils at 62°.

Free sulphuric acid can be immediately detected by a solution of any salt of baryta. When one of these is added, the heavy white sulphate of baryta, insoluble in all acids and alkalis, is at once thrown down.

In order to ascertain the proportion of sulphuric acid really present in any commercial specimen, it may be moderately diluted with water, and while warm pure anhydrous carbonate of soda added until the solution is quite neutral. Every 53 parts of carbonate of soda correspond to 40 of real acid, these being respectively the atomic weights of these two substances.

5. HYPOSULPHUROUS ACID.

It is unnecessary to detail here the proportion of this compound. It cannot exist permanently in a free state. Nevertheless, it forms salts, and the hyposulphites are more permanent than the sulphites. A curious property of the neutral hyposulphites is that they can dissolve in large quantity chloride of silver, and that they form with that substance a liquid of an exceedingly sweet taste.

6. HYPOSULPHURIC ACID.

This is another compound of little interest save to the professed chemist, and may here be safely omitted.

7. CHLORIDE OF SULPHUR.

This consists of one atom of chlorine united with two of sulphur ; is prepared by simple admixture of chlorine and sulphur vapour, and is a volatile reddish liquid, of no practical importance.

8. BROMIDE OF SULPHUR.

This, like the preceding, is a heavy reddish liquor.

9. IODIDE OF SULPHUR.

This is formed by heating gently together four parts of iodine with one of sulphur. The result is a black crystallisable compound. It is used as an ointment in various skin diseases.

* * *

VIII. Selenium.

This elementary substance was discovered by Berzelius, and accompanies, but in exceedingly minute quantities, metallic sulphurets. Neither it nor its compounds are put to any use in the arts, and it only requires here a very short notice.

Selenium is a solid, having a dark brown colour and a metallic lustre. It boils at 650°, giving off a vapour of a yellow colour like that of sulphur, a substance which in its combination it is also analogous to.

Thus it forms with hydrogen a gas, seleniuretted hydrogen, having a very fetid smell, being irrespirable, and precipitating from metallic solutions selenurets. These are commonly black, but that of zinc is white.

Unlike sulphur, selenium forms a compound with oxygen having only one atom of that element. It is a gas remarkable for its smell, which resembles that of horse-radish. It also forms selenious and selenic acids, both which are very analogous to sulphurous and sulphuric ones.

Its atomic weight is 39, and its symbol Se.

IX. Phosphorus.

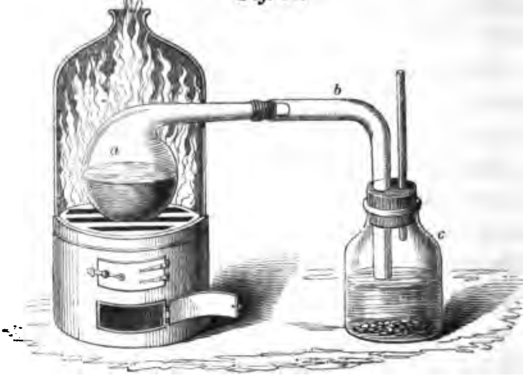
The history of the discovery of this element is curious. It was found out by accident by Brandt, an alchemist of Hamburgh, who was in search of the philosophers' stone. "Like his *confrères*," says Dumas, "he did not find that which he sought, but, more happy, he found that which he did not seek." For some time he kept his process secret; but it was discovered by Kunkel, whose phosphorus was sold under his name. The process employed was tedious, and for a long time phosphorus was rare and dear. When Scheele discovered that the bones of animals contain phosphorus, a new source of it was brought to light, from which it can now be abundantly obtained.

The name phosphorus is given to it on account of its being luminous in the dark; and is derived from *φῶs light, φέρω to carry.*

Phosphorus occurs abundantly in nature. It is a necessary constituent of bones, and therefore of the food of animals, and therefore of edible plants, and therefore of the soil. Indeed, one of the difficulties that the farmer has to contend with is to keep up the supply of soluble phosphates in severely cropped land.

To obtain phosphorus bones are burnt to expel the animal matter. The phosphate of lime that is then left behind is ground down to powder. To this water and sulphuric acid are added, insoluble sulphate of lime is formed, and the phosphoric acid, which remains dissolved in the water, is set free. The liquid is evaporated to the consistence of a syrup, mixed with a quantity of charcoal, and the mass is completely dried at a temperature below that of a red heat. It is then powdered and introduced into the retort *a* (*fig.* 31.), which is placed in a furnace. A copper tube, *b*, is adapted to the neck of the

Fig. 31.



retort, the other end of which dips into the water in the bottle *c*. As the retort becomes heated the carbon of the charcoal unites with the oxygen of the phosphoric acid, forming carbonic acid, and the phosphorus becomes free; and, being volatilised by the high temperature, passes along *b*, where it is condensed into a liquid form whilst under the water. The phosphorus is again melted under the water and poured into glass tubes, where it allowed to solidify. It then forms the sticks of phosphorus sold in the shops.

Phosphorus, when pure, is transparent, and nearly colourless. At common temperatures, it is a soft solid, of specific gravity 1·7; it is easily cut with a knife, the cut surface showing a waxy lustre; at 108° it melts, and vaporises at 550°. It is soluble in naphtha, oils, &c.

Phosphorus is remarkably inflammable. Exposed to the air, at common temperatures, it slowly burns; emits a white vapour that has a strong smell of garlic; is luminous in the dark, and it is gradually consumed, and consequently, it requires to be kept under water.

If, when exposed to the air, the slightest heat be applied to it, as even that communicated to it by slight friction, it kindles, emits a splendid white light, and gives out intense heat. If it be burned in pure oxygen the combination is even more vivid.

Its atomic weight is 16, and its symbol P. The combining measure of its vapour is □.

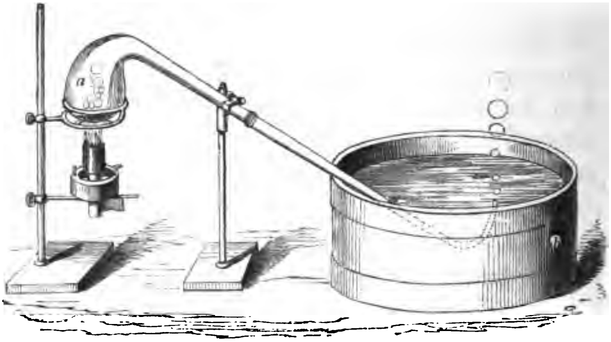
The compounds of phosphorus described in this section are as follow: —

1. *Phosphuretted hydrogen* -	Phosphorus 1 atom	16+Hydrogen	3 atoms	3=19	$P+3H$.
2. *Oxide of phosphorus* -	Phosphorus 3 atoms	48+Oxygen -	1 atom	8=56	$3P+O$.
3. *Hypophosphorous acid* -	Phosphorus 2 atoms	32+Oxygen -	1 atom	8=40	$2P+O$.
4. *Phosphorous acid* - -	Phosphorus 2 atoms	32+Oxygen -	3 atoms	24=56	$2P+3O$.
5. *Phosphoric acid* - -	Phosphorus 1 atom	16+Oxygen -	5 atoms	40=56	$P+5O$.
6. *Pyrophosphoric acid* -					
7. *Metaphosphoric acid* -					
8. *Iodide of phosphorus* -	Various.				
9. *Sulphuret of phosphorus*	do.				

1. PHOSPHURETTED HYDROGEN.

There are probably two compounds of phosphorus and hydrogen, but our knowledge is certain with respect to the gaseous one only. This gaseous one was discovered by Davy in 1812. It may be formed, either by heating phosphorus in lime water, in which case the water is decomposed, the oxygen uniting with part of the phosphorus, and the hydrogen with the other, the latter union forming the compound under consideration; or by heating hydrated phosphorus, in which case the water is also decomposed. As the gases obtained in these two methods have different proportions, Kane has proposed to call the one procured by the first process, A gas, and the other, B gas.

Fig. 32.



If the A gas evolved from the retort *a* (*fig.* 32.) be allowed to bubble through the water of the pneumatic trough *b*, each bubble, on reaching the surface, bursts into flame, and becomes converted into a wreath of white smoke, that consists of phosphoric acid and water. As these rings ascend in the air, they widen, and present a curious combination of movements. Each ring, in fact, consists of an amazing quantity of little rings, revolving with great rapidity on their axis, at right angles to the direction of the large one. If these bubbles are received into oxygen, the combustion is very brilliant.

The B gas does not spontaneously inflame, although, if ignited, it burns in the same manner.

In other respects, the two are identical. They are colourless, transparent gases, having a most offensive smell; they precipitate most metallic salts, giving insoluble phosphurets; and they consist of one atom of phosphorus and three of hydrogen.

The difference in the two, with regard to their inflammability, is owing to the presence of a little foreign matter. Thus, the addition of a very small quantity of ether destroys the spontaneous inflammability of A gas, and, on the other hand, a little vapour of nitrous acid converts B into A, and makes it spontaneously ignite. Graham suggests that, in the preparation of the gas from phosphorus and lime water, a compound of oxygen and phosphorus may be formed, the same as that formed by nitrous acid with B, and which acts as a match.

2. OXIDE OF PHOSPHORUS.

To obtain this, phosphorus is dissolved in boiling water, and then a stream of oxygen gas is thrown upon it. Combustion ensues, the phosphorus burns with great brilliancy, and a number of red particles fall to the bottom. It is a compound of three atoms of phosphorus and one of oxygen.

3. HYPOPHOSPHOROUS ACID.

This, like the preceding, is a substance of little consequence. It forms salts, nearly all of which are soluble in water.

4. PHOSPHOROUS ACID.

This acid is the principal product of the slow combustion of phosphorus. In order to obtain it in this manner, the phosphorus must be in excess. Thus prepared, phoshporous acid is a soft, white powder.

It is also obtained in the liquid form. To do so, a current of chlorine is forced upon phosphorus placed in water, upon which chloride of phosphorus is formed; this substance is immediately decomposed by the water, and hydrochloric acid and phosphorous acid are produced. By evaporating this mixture to the consistence of a syrup, the hydrochloric acid passes off as a gas, leaving hydrated phosphorous acid behind. If this be heated, the water is not driven off; but it is decomposed, and resolved into phosphoric acid and the kind of phosphuretted hydrogen that is not spontaneously inflammable.

The solution of phosphorous acid has a strong affinity for oxygen, and, by the aid of heat, reduces to the metallic state salts of mercury, silver, gold, and platina.

5. Phosphoric Acid.

This acid may be prepared by oxidating phosphorus by means of nitric acid. The following table shows the action that takes place.

Ingredients used.			Results.	
5 atoms nitric acid -	5 atoms nitrogen	70	5 atoms binoxide of nitrogen	150
	10 atoms oxygen	80		
	15 atoms oxygen	120	6 atoms phosphoric acid -	216
6 atoms phosphorus		96		

Phosphoric acid is colourless, very sour, reddens litmus strongly, and neutralises acids; it does not, however, destroy the texture of the skin. It has a very strong affinity for water, and when brought into contact with that substance combines almost explosively. It forms with it three distinct compounds or phosphates of water, as follows:—

Monobasic phosphate of water $=(P+5O)+HO$.
Bibasic phosphate of water - $=(P+5O)+2HO$.
Tribasic phosphate of water - $=(P+5O)+3HO$.

Not only does the phosphoric acid combine in these different proportions with water, but these phosphates of water may be decomposed and the place of the water taken by metallic oxides. When the monobasic phosphate of water is so decomposed, say by means of soda, a phosphate of soda with one atom of acid and one of base is formed; when the bibasic phosphate of water is so decomposed, then two atoms of the soda take the place of the two of water, and join with the one of acid; and when the tribasic phosphate of water is thus decomposed, then three atoms of soda join with one of base.

This was formerly explained by assuming that there were three isomeric phosphoric acids, and which were named as above; but the more recent views of Professor Graham are not only the most probable, but they have a very important bearing upon the constitution of salts in general.

6. IODIDE OF PHOSPHORUS.

When iodine and phosphorus are warmed together they combine and form, according to the proportions used, three iodides, an orange, a grey and a black one.

7. SULPHURET OF PHOSPHORUS.

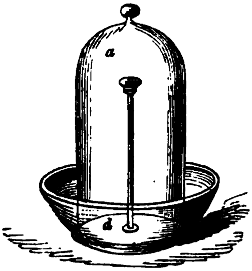
In like manner, when sulphur and phosphorus are mixed together, they form one or more unions. Sulphuret of phosphorus is very inflammable, and is used in the phosphorus match-boxes.

X. Azote, or Nitrogen.

This element occurs very abundantly in nature. It constitutes four-fifths of the atmosphere; occurs in the soil in the form of nitrates; is essentially present in animal and vegetable structures, and is, by the decay of these, also supplied to the soil, and to the air as ammonia, which, as we shall soon see, is a nitrogenous compound. The name of nitrogen is derived from its being an element of nitric acid; and azote from *ἀ, ζωὴ, life,* in consequence of animals placed in it soon dying. It is not, however, that animals so treated die because they breathe the nitrogen, but because they do *not* breathe oxygen.

Nitrogen is sometimes obtained from flesh, but its most convenient source is the atmosphere. By depriving this compound of its oxygen, a sufficiently pure nitrogen is obtained. This is best done by burning a piece of phosphorus in a cup, *d* (*fig.* 33.), floating on water, and inverting a bell glass, *a*, over it. The phosphorus and oxygen unite, and form phosphoric acid, which is absorbed by the water. As the oxygen gradually disappears, its place is taken by water, and about one-fifth of the bell glass will be occupied by that fluid. The carbonic acid, ammonia, and other constituents of the atmosphere are separated from the nitrogen by washing.

Fig. 33.



Nitrogen is remarkable for its negative properties. It is without colour, taste, or smell, is not combustible, and does not support combustion or respiration. It is rather lighter than atmospheric air.

Its atomic weight is 14, its symbol N, and its combining measure ◻◻.

The compounds of nitrogen mentioned or described in this section are as follows:—

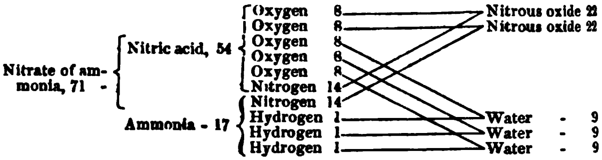
			Formula.
1. *Azuret of hydrogen, or ammonia*	Nitrogen 1 atom 14+Hydrogen 3 atoms	3=17	$N+3H$.
2. *Nitrous oxide* -	Nitrogen 1 atom 14+Oxygen - 1 atom	8=22	$N+O$.
3. *Nitric oxide* -	Nitrogen 1 atom 14+Oxygen - 2 atoms	16=30	$N+2O$.
4. *Hyponitrous acid* -	Nitrogen 1 atom 14+Oxygen - 3 atoms	24=38	$N+3O$.
5. *Nitrous acid* -	Nitrogen 1 atom 14+Oxygen - 4 atoms	32=46	$N+4O$.
6. *Nitric acid* -	Nitrogen 1 atom 14+Oxygen - 5 atoms	40=54	$N+5O$.

1. AZURET OF HYDROGEN, OR AMMONIA.

This compound was probably not created along with other material compounds, but was only called into being, on anything like an extensive scale, after organic structures had lived and rotted. Its consideration is therefore deferred until we have, in Book III., organic chemistry brought before our notice.

2. NITROUS OXIDE.

Nitrous oxide, or the protoxide of nitrogen, was discovered by Priestley, who named it dephlogisticated air. It is best prepared by heating nitrate of ammonia up to 350°, at which temperature this salt is decomposed, and resolved into nitrous oxide and water. The nature of this very curious change is shown by the following diagram:—

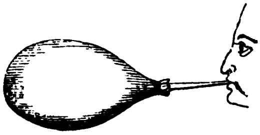


Nitrous oxide is a colourless and transparent gas. It is heavier than air, having a specific gravity of 1·5. Its odour and taste are agreeable. It is not combustible, but is a very powerful supporter of combustion; in this respect almost equalling oxygen. Water dissolves about three-fourths of its volume of it.

As nitrous oxide is an invisible gas that supports combustion, it is apt to be mistaken for oxygen. It may be distinguished from that element by its solubility in water, and by sulphur burning in it with a rose tint, and not with a blue or purple flame as in oxygen.

The effects of nitrous oxide, when inhaled, upon the nervous system are very remarkable. If breathed from a waterproof bag (as in *fig.* 34.) it produces a species of temporary intoxication. "I have administered," wrote Pereira, "this gas to about one hundred persons, and have observed that, after the respiration of it from a bladder for a few seconds, it usually causes frequent and deep inspirations, blueness of the lips and countenance, and indisposition to part with the inhaling tube, and a temporary delirium, which subsides in the course of three or four minutes.

Fig. 34.



The sensations are usually pleasing. The delirium manifests itself differently in different individuals, as in some by dancing, in others by fighting, &c." After the excitement has worn off, little or no lassitude is felt.

3. NITRIC OXIDE.

Nitric oxide, or the deutoxide of nitrogen, is a gas prepared by deoxidating nitric acid. Nitric acid contains five atoms of oxygen, nitric oxide two; if, then, we extract from an atom of the former three atoms of oxygen, we obtain the latter. This is done by adding to nitric acid metallic copper. Brisk effervescence begins without the aid of heat, nitrate of the oxide of copper is formed, and nitric acid gas given off. The chemical change that occurs is as follows:—one portion of the nitric acid is decomposed; a portion of its oxygen oxidates the copper, and the oxide thus formed unites with the other portion of nitric acid; and the other portion of oxygen unites with the nitrogen to form the compound in question.

Nitric oxide is a colourless, transparent gas, a little heavier than air, scarcely soluble in water, irrespirable, incombustible, and not a supporter of ordinary combustion.

Its most remarkable property is its tendency to unite with oxygen when the oxygen is uncombined. It cannot take oxygen from any other substance; but if it is mixed with it, or with a mixture of gases of which oxygen is one, it unites with it, forming the deep ruddy fumes of nitrous acid. Thus nitric acid forms a very ready test for the presence of free oxygen. Nitric oxide combines with both salts and acids. These compounds will be afterwards named.

4. Hyponitrous Acid.

This is formed by mixing, in proper proportions, nitric oxide and oxygen, and surrounding the vessel with a freezing mixture. They unite to form a volatile green liquid.

5. Nitrous Acid.

Nitrous acid presents itself in the form of deep red fumes, and is produced by simple admixture of its elements, the oxygen being in excess. By a cold of about zero it may be rendered liquid. This fluid nitrous acid at this temperature is of an orange-yellow colour; at 40° it solidifies, and then becomes white; at 82° it boils, and its vapour is then ruddy red; but if it be heated to 212° becomes black. These changes of colour are very remarkable.

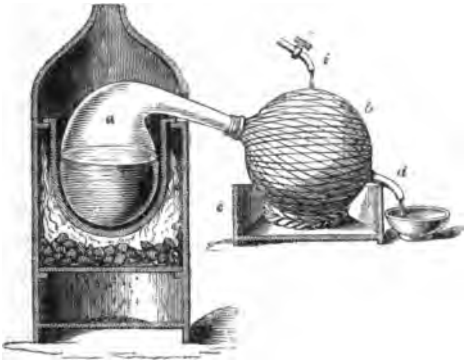
It is now believed to form with bases salts which are termed nitrites.

6. Nitric Acid.

This important acid was known in the seventh century to Geber, who named it *solutive water*. It subsequently received various names, as *Glauber's spirit of nitre*, *aquafortis*, &c. It occurs combined with lime, potash, soda and magnesia, is often produced in the atmosphere by the action of electricity upon its elements, and is present in many vegetables. Anhydrous nitric acid is composed of one atom of nitrogen and five of oxygen, but it is nearly always, and in practice always, combined with water.

On a small scale hydrated nitric acid* is obtained from nitrate of potassa by the action of sulphuric acid. Equal weights of the two ingredients are placed in the glass retort *a* (*fig.* 35.), to which heat is applied. Nitric acid soon begins to distil over, and is condensed in the receiver *b*. In order to render this condensation more complete, the receiver is surrounded by a net, and placed in a trough, *c*, and a stream of cold water is made continually to flow over it from the pipe *i*.

Fig. 35.



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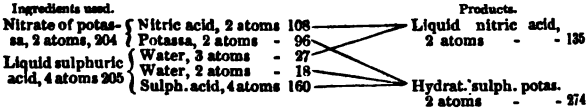
* In the following pages we shall term the acid simply "nitric acid."

This cold water spreads over all parts of the network, and flows out by the end pipe *d*. This apparatus is suited for the distillation of other volatile liquids.

The reaction that takes place in this process may be exhibited thus:—


Ingredients used.
Nitrate of potassa, 2 atoms, 204
Liquid sulphuric acid, 4 atoms 205
Nitric acid, 2 atoms 108
Potassa, 2 atoms - 96
Water, 3 atoms - 27
Water, 2 atoms - 18
Sulph. acid, 4 atoms 160
Products.
Liquid nitric acid, 2 atoms - - 135
Hydrat. sulph. potas. 2 atoms - - 274

In practice, however, nitric acid is not prepared on the small scale in glass retorts, but in large quantities by the manufacturers in iron vessels. The salt employed by them is now almost invariably the nitrate of soda, it being much cheaper than the saltpetre. The nitric acid of commerce is very far from being pure. It generally contains chlorine (from the nitrate used having common salt mixed with it), sulphuric acid, iron from the retort, and nitrous acid. By distillation it may be rendered tolerably pure, the chlorine coming off with the portion that comes over first, and the salt and iron remaining to the last. The acid that is distilled over in the middle is the purified kind.

The strongest nitric acid that can be obtained consists of one atom of acid and one of water. This has a gravity of 1·52, and is very rarely procured. In general the strong acid consists of three atoms of water and two of acid. This has a specific gravity of 1·50.

Nitric acid is, when pure, without colour. When exposed to the air it fumes, and if exposed to direct solar light it is decomposed, being converted into nitrous acid and oxygen. This decomposition is produced at once by passing the acid in a state of vapour through a red-hot porcelain tube. Several of the non-metallic combustibles, as charcoal, phosphorus, sugar, &c., and also many metals, rapidly decompose it and appropriate its oxygen. Hence, when it is wished to oxidise a substance, nitric acid is very frequently employed.

Applied to animal substances nitric acid forms compounds of a yellow colour, as is exemplified if a drop of it fall upon the hand. On account of this property it is employed for stamping yellow patterns on woollen table cloths.

The presence of nitric acid is easily detected. It is known by its producing red fumes when brought into contact with copper. These red fumes are of course those of nitrous acid. Passed into a solution of protosulphate of oxide of iron it forms a peculiar olive-brown coloured liquor. This is owing to one portion of the protosulphate reducing the acid to the state of nitric oxide, which then combines with the remainder of the salt of iron. Nitric acid also strikes a deep red colour with morphia.

Nitric acid is extensively used in the arts.

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XI. BORON.

Boron was first obtained by Davy, who procured it by exposing boracic acid to the action of a voltaic
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battery; but it is now prepared by the action of potassium upon fluoborate of potash. This last-mentioned salt consists of fluoride of boron united to fluoride of potassium; and when the potassium is added to it all the fluorine passes to it, leaving the boron free.

Boron is a dark olive substance, devoid of taste or smell. It does not conduct electricity, is not soluble in water, and is not acted upon by any of the acids except nitric, which parts with some of its oxygen to it, and converts it into boracic acid.

Its atomic weight is about 11, and its symbol B.

The compounds of boron mentioned in this section are as follow:—

					Formula.
1. *Boracic acid* -	Boron 1 atom=11	+Oxygen -	3 atoms	24= 35	$B+3O$.
2. *Fluoride of boron*	Boron 1 atom=11	+Fluorine	3 atoms	56= 67	$B+3F$.
3. *Chloride of boron*	Boron 1 atom=11	+Chlorine	3 atoms	106=117	$B+3Cl$.

1. Boracic Acid.

Boracic acid is found in certain springs in India combined with soda, and called in commerce tincal. It also occurs free in the lagoons of Tuscany, and is a constituent of two or three minerals. It is usually prepared by adding sulphuric acid to a hot solution of borate of soda. The sulphuric acid unites with the soda, and the boracic acid falls down in shiny, scaly crystals. As, however, it still retains a little sulphuric acid and sulphate of soda, it requires to be purified by a second solution and crystallisation.

Boracic acid thus prepared contains water, every atom of it being associated with three atoms of that fluid. This hydrated acid is very much more soluble in hot water than in cold. Boiling alcohol also freely dissolves it; and this solution when set on fire burns with a beautiful green flame, a property which is quite characteristic of this acid. Boracic acid is a very feeble acid, it does not redden litmus; a strong solution of it browns turmeric like an alkali, and its salts (called borates) are very easily decomposed.

If the hydrated acid be heated the water is driven off and the acid melts. On cooling it forms a hard, transparent, vitreous substance, called glacial boracic acid. It is extremely fusible, and communicates that property to the substances with which it unites, and is hence much used as a flux.

2. FLUORIDE OF BORON.

This is sometimes called fluoboric acid. It may be prepared by decomposing fluor spar by means of vitrified boracic acid. Oxygen and calcium unite together and form lime, which allies itself with the boracic acid, and forms borate of lime, and a part of the fluorine and of the boron unite and form the compound in question.

It is a gas that is very rapidly absorbed and decomposed by water, hydrofluoboric acid being generated.

When exposed to the air, if moisture be present, dense white fumes are produced, owing to its being decomposed and boracic and hydrochloric acids being formed.

3. Chloride of Boron.

This is a transparent colourless gas, remarkable for its great specific gravity, which is 4076.

XII. Silicon.

This element is very widely distributed in nature, and forms perhaps one sixth of the crust of the globe. It is, however, always united with oxygen, forming silicic acid, or in popular language sand.* It enters into the structure of many vegetables, but never occurs in animals.

Silicon may be obtained by heating dry silicic acid along with potassium. The potassium unites with the oxygen and forms potassa, which unites with a portion of the silicic acid and becomes silicate of potassa. The free silicon may then be obtained pure by washing with water.

Silicon is a dark brown substance, without any metallic lustre, and it does not conduct electricity. It is insoluble in water, not readily acted upon by liquid re-agents, and if strongly heated in air or oxygen combines with the oxygen, forming silicic acid.

Its atomic weight is 22, and its symbol Si.

---

* Besides common sand, which is disintegrated silicious rock, silicic acid is the principal ingredient in quartz, flints, agates, and many other minerals.

The compounds of silicon noticed in this chapter are as follow: —

		Formula.
1. *Silicic acid* - -	Silicon 1 atom 22+Oxygen - 3 atoms 24= 46	$Si+3O$.
2. *Chloride of silicon*	Silicon 1 atom 22+Chlorine 3 atoms 106=129	$Si+3Cl$.
3. *Fluoride of silicon*	Silicon 1 atom 22+Fluorine 3 atoms 56=78	$Si+3Fl$.
4. *Hydrofluosilicic acid*	A double fluoride of hydrogen and silicon.	

1. SILICIC ACID.

This exists pure in the form of quartz, and, combined with other substances, in great abundance in many other minerals. It is remarkably hard, and if it be desired to obtain it in powder, it is necessary to heat it to redness, and throw it in that state into water. It may then be powdered in an agate mortar. It may be procured in a state of much more minute division by melting in a platinum crucible equal weights of carbonate of soda and potassa, and then adding powdered flint in small quantities at a time. Carbonic acid is given off, and the silica unites with the alkalis, and silicates of soda and potassa are formed. These are then dissolved in water, and to the solution hydrochloric acid is added; this unites with the bases, and hydrated silicic acid is precipitated. This, when dried, is a white powder; still, however, gritty to the feel. Another mode of obtaining silica in a still more minute powder is to bring fluoride of silicon in contact with water; upon which a portion of the water is decomposed, and fluoride of hydrogen and silica are produced.

Silica has a specific gravity of 2·6. It is infusible by the most powerful wind furnace, but yields before the oxy-hydrogen blowpipe. Silica once dried is quite insoluble in water; but in its hydrated condition it is soluble to a certain extent, and many waters contain traces of it, and it is considered probable (Kane) that most of the silica of this world was originally in a soluble state.

Silica is not sour to the taste, neither does it affect vegetable colours; but it unites readily with bases forming silicates. As will be more immediately seen in Book II., these silicates occur very abundantly in the structure of the globe.

Silica is very much used in the arts. Glass, porcelain, stone, and china, are all artificial silicates.

2. Chloride of Silicon.

This compound is not of importance. It is a colourless, volatile liquid.

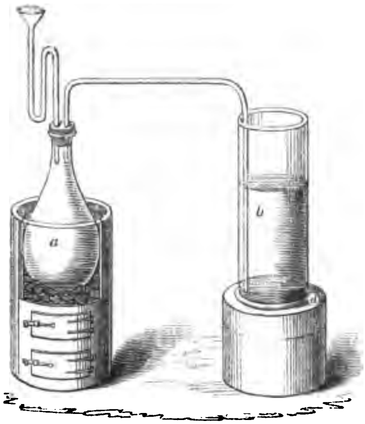
3. Fluoride of Silicon.

In order to prepare this, fluoride of calcium (fluor spar) sand and sulphuric acid are mixed together and heated. The oxygen of the silica unites with calcium, forming lime, to which the sulphuric acid attaches itself; while the silicon and fluorine combine and form the compound now under our notice. It is a colourless and transparent gas, which extinguishes flame. The specific gravity is 3·6. It must be collected over mercury, as it is very powerfully acted upon by water. When it mixes with the air, white fumes are produced, owing to silica being formed by the watery vapour being decomposed.

4. Hydrofluosilicic Acid.

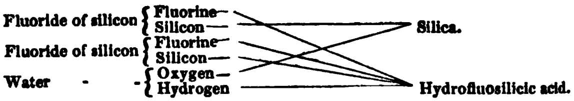
Skilful management is required to make this compound. The fluor spar is introduced into the glass vessel *a* (*fig.* 36.), which is placed in a sand bath; sulphuric acid is poured into it by means of the syphon funnel, and the gaseous fluoride of silicon that is evolved is conducted into the vessel *b*, which contains water. If, however, the gas opened directly into the water, so much silica would be immediately formed as to instantly stop up the tube; a quantity of mercury, therefore, *d*, is placed at the bottom, and the end of the tube made to dip into it. In consequence of this arrangement the gas bubble does not reach the water until quite separated from the tube, and when it emerges from the mercury becomes converted into

Fig. 36.



a little bladder of silica. When the water has become thick, the process is stopped, the fluid strained, to remove the silica, and the fluid then obtained is a solution of the acid under consideration.

The changes that take place in this process are exhibited in the following table: —



This acid is valuable as a test for potassa and baryta.

CHAP. VII.

THE METALS, AND THEIR BINARY COMPOUNDS WITH NON-METALLIC ELEMENTS.

A METAL is an elementary substance, characterised by conducting heat and electricity, and by reflecting light. This last property gives metals a peculiar lustre, named the metallic.

All the metals, when pure, are tasteless, although their salts have often very strong and disagreeable tastes. In colour they vary a good deal; silver is white, gold yellow, copper red, and the remainder bluish white or grey. The different metals, too, are of very different specific gravities. Platinum, for example, is twenty-one times heavier than water, and potassium is lighter than that fluid, and consequently swims upon its surface. The metals commonly used have generally a specific gravity of about six or eight times that of water.

At ordinary temperatures all the metals save one, namely mercury, are solid, but by heat they can all be rendered fluid; but they require a very wide range of temperature to liquefy them. Potassium, for example, melts at a heat below that of boiling water, tin at less than a red heat, while platinum only yields under the intensity of the oxy-hydrogen blowpipe.

Most of the metals are fixed in the furnace; but seven, mercury, zinc, cadmium, arsenic, tellurium, potassium, and sodium may be volatilised.

The malleability and ductility of many of the metals are put to very useful purposes in the arts. A malleable metal is one that may be rolled or beaten into thin leaves, and a ductile, one that may be drawn into wires. Of all metals gold is the most malleable; and gold leaf, as it is called, may be procured of the almost incredible size of $\frac{1}{282,000}$ in thickness or rather thinness. The most ductile metal is platinum, the wire of which may be made the $\frac{1}{30,000}$ of an inch in diameter. Some of the metals are neither malleable nor ductile, but this as well as other peculiarities in this respect will be considered under their proper heads.

Generally speaking metallic bodies have a strong affinity for oxygen, and if exposed to the air, especially if the air be damp, unite with oxygen and become converted into oxides. The degree of affinity however that metals show for oxygen varies very much. That of potassium and others is so strong, that water is decomposed by them with great effervescence, even at a temperature of 32°. Others, again, as manganese, &c., can only appropriate the oxygen of water at a temperature of 212°; a third class, including iron and tin, do not decompose water save at a red heat, or unless strong acids are present; while others, again, as silver, gold and others, do not decompose water under any circumstances.

Many of the metals combine with oxygen in more than one proportion. The protoxides (as the suboxides) are almost invariably strong bases, uniting with acids to form neutral salts. The protoxide of arsenic is an exception to this, as it has a decidedly acid action; and hence in some works upon chemistry this element is ranked among the non-metallic substances. The peroxides of the metals are either neutral or acid; and when, as is sometimes the case, the degree of oxidation becomes very great, these compounds become highly acid, and by uniting with bases form salts.

Sometimes, indeed, metallic oxides unite together very much after the fashion of acids and bases, and produce compounds that are very similar to neutral salts.

Metals unite with chlorine, forming chlorides, even more readily than they do with oxygen. They also readily unite with iodine and bromine, thereby giving rise to iodides and bromides.

Their tendency to unite with sulphur is also very great; and the sulphurets so formed are opaque, brittle solids, many of which have a metallic lustre. The protosulphurets in general may be exposed to an intense heat without being decomposed; but those that contain more than one atom of sulphur lose part part of it when strongly heated, and all of them can be decomposed by the combined action of air (or oxygen) and heat.

It is necessary also to state that hydrogen combines with zinc and another metal or two; that many phosphurets occur in nature, or may be formed; and that there is one important carburet, the carburet of iron or steel.

The metals are found in nature in very different states. Some are found pure, or native as it is termed; others are mixed, or alloyed with other metals. Many chlorides, iodides and phosphurets are also found as minerals; but the greatest portion of natural metallic compounds, or *ores* as they are called, are either oxides or sulphurets.

The mode of extracting the different metals from their ores belongs to Book IV.; but some general observations upon the separating of the pure metal from oxides and sulphurets will be proper here.

Metallic oxides suffer *reduction*, as it is termed, in various ways.

1st. By the application of heat alone. By this oxides of gold, silver, &c., may be separated into oxygen, which is given off as gas, and the metal.

2nd. By the galvanic battery. This is the most effectual deoxidiser of any.

3rd. By the united agency of heat and combustible matter. In this manner, if we transmit a stream of hydrogen over oxide of iron heated to redness in a porcelain tube, the oxygen and hydrogen unite to form water, and the pure metal is separated. Carbonaceous matter is very efficacious in a similar manner, and potassium may be obtained from potassa by heating this latter substance along with charcoal. In this instance, of course, besides potassium, carbonic acid is formed. A plan similar to this is generally used in the arts, wood or coal being the combustible employed.

4th. By the action of deoxidising agents on metallic solutions. Thus, if phosphorous acid be added to a liquid containing solution of oxide of mercury, the oxygen unites with the phosphorous acid, and the two form phosphoric acid, the metallic mercury being set free.

The metal, in the case of some metallic sulphurets, may sometimes be separated by melting the one along with a metal that has a stronger affinity for sulphur. Thus, metallic arsenic may be obtained by melting its sulphuret with iron. Such a plan is only applicable on a small scale.

The metals are now usually arranged in accordance with their affinity for oxygen, as is indicated by the facility with which they decompose water.

A.	Potassium Sodium Lithium Barium Strontium Calcium Magnesium	This first class of metals consists of those that decompose water with effervescence, at 32°.
B.	Aluminum Glucinum Thorium Ittrium Zirconium Lanthanum Cerium Manganese	This second class of metals consists of those that decompose water very far below a red heat, but not below about 212°.
C.	Iron Nickel Cobalt Zinc Cadmium Tin Chromium Vanadium	This third class consists of metals which do not decompose water, except at a red heat, or at common temperatures when in contact with strong acids.
D.	Tungsten Molybdenum Osmium Columbium Titanium Arsenic Antimony Tellurium Uranium	This fourth class consists of metals which decompose water at a red heat, but which do not decompose it at common temperatures, even when in contact with strong acids.
E.	Copper Lead Bismuth	The fifth class consists of metals which decompose, although very feebly, water at a red heat, but the oxides of which are not reducible to the metallic state by heat alone.
F.	Silver Mercury Gold Palladium Platinum Rhodium Iridium	The sixth class consists of metals which do not decompose water under any circumstances. Their oxides are reducible by heat alone.

The old nomenclature and arrangement of the metals, although never systematically used now, is still frequently referred to, and a knowledge of it is indispensable to the understanding of old chemical works.

Those metals that do not oxidate on exposure to the air, and whose oxides are reducible by heat alone (gold, silver, &c.), were called the noble or perfect metals; all the other were denominated imperfect metals. Only the oxides of the first two of the above enumerated classes (A. p. 163.) were known to the antients, and potassa and soda were termed alkalis. To these they joined ammonia (see Book III.), because these three substances agreed in being soluble in water, and neutralising the strongest acids; in changing vegetable yellows to browns, and reds and blues to greens. Ammonia was sometimes called the volatile alkali, and the other two the fixed alkalis. Ammonia, being obtained from the putrefaction of animal matter, was often also called the animal alkali; soda, being procured from the earth, was the mineral alkali; and potassa was the vegetable alkali, because its usual source was, as it is, the ashes of plants.

Magnesia and alumina were termed earths, and silica was classed along with them. Lime, barytes, and strontia were called alkaline earths, because they were soluble in water, formed salts with acids, and turned brown, yellow vegetable colours. They were distinguished from the alkalis proper by the insolubility of their carbonates.

We now proceed to the consideration of the particular metals, and their binary compounds with the non-metallic elements.

* * *

(A.) I. POTASSIUM.

Potassium occurs very abundantly in nature, but never in a free state, being always combined with oxygen (forming potassa). It is a constituent of a great many minerals, and is essentially present in all fertile soils. It is a necessary ingredient in the structure of animals and plants, and is returned to the soil by the decomposition of these.

Potassium was discovered by Davy in 1807. He obtained it by subjecting potassa to the action of a powerful voltaic battery, upon which the oxygen was detected at one pole, and the potassium at the other. The usual mode of procuring it now is by the deoxidising effect of heat and charcoal upon potassa.

Potassium is a bluish white metal, hard and brittle at 32°, and soft and malleable at 50°; fluid at 150°, and volatilised at a temperature below a red heat. At 60° its density is ·860, so that it is considerably lighter than water. It is quite opaque, and conducts heat and electricity well.

Its most prominent chemical property is its very strong affinity for oxygen. In contact with the air or with gas, or fluid that contains oxygen, it is rapidly oxidated; and must therefore be kept in some fluid, as naphtha, of which this element is not an ingredient. The moment that it touches water it decomposes it, and so much heat is disengaged, that the potassium takes fire on the surface of the water. The greater portion of the potassium combines with the oxygen of the water; but a portion unites with the liberated hydrogen, forming a hydruret of potassium, which also takes fire. If the potassium be plunged right under the surface of the water, there is violent reaction, but no flame; and pure hydrogen gas is evolved.

Its symbol is K, and its atomic weight 39.

The compounds of potassium mentioned in this section are—

		Formula.
1. *Hydruret of potassium* -	Composition uncertain.	
2. *Protoxide of potassium* - -	Potassium 1 atom 39+Oxygen - 1 atom 8= 47	$K+O$.
3. *Peroxide of potassium* - -	Potassium 1 atom 39+Oxygen - 3 atoms 24= 63	$K+3O$.
4. *Chloride of potassium* - -	Potassium 1 atom 39+Chlorine 1 atom 35= 74	$K+Cl$.
5. *Iodide of potassium* - -	Potassium 1 atom 39+Iodine - 1 atom 126=165	$K+I$.
6. *Sulphuret of potassium.*	There are five sulphurets.	

1. Hydruret of Potassium.

Two compounds (perhaps of the nature of alloys) are formed by hydrogen and potassium; the one a gas that spontaneously burns in air, and the other a solid which does not spontaneously burn. Neither of them is of importance.

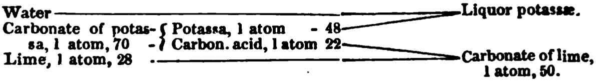
2. Protoxide of Potassium, or Potassa.

This compound is always formed when potassium is put into water, or into air or pure oxygen. In the former of these cases it is obtained as a hydrate, in the latter in an anhydrous state.

Anhydrous potassa is a white, solid substance, intensely caustic, and having a strong affinity for water. Indeed, the causticity depends upon its destroying any structure to which it is applied, by abstracting from it all its water. In the arts it always exists as the —

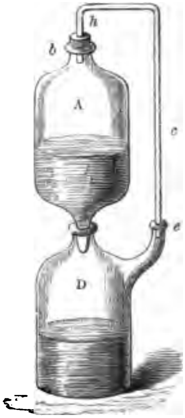
Hydrate of Potassa.

A solution of this, called *liquor potassæ,* is made by adding fresh burnt lime to a solution of carbonate of potassa. The lime abstracts carbonic acid from the carbonate of potassa, and the potassa, thus set free, dissolves in the water.



The carbonate of lime is separated by filtering. Notwithstanding the carbonate of potassa parts with its carbonic acid to the lime, it has a strong affinity for that acid; and as carbonic acid always exists in the air, it is necessary to filter its solution in vessels containing as little air as possible. This is effected by using the apparatus, *fig.* 37. A and D are two vessels of equal capacity, the bottom of A being a funnel opening into D. *c* is a glass tube, one end opening into each at *h* and *e*. By means of a cork the joinings are all made air-tight. Into the throat of the funnel pipe some coarse linen is placed, the solution is then poured into A at *b*, and the tube adjusted. In this manner there is not any exposure

Fig. 37.



to more air than was contained in the vessel at the beginning of the process.

Solution of potassa thus obtained is very caustic, and highly alkaline. It is a limpid, colourless fluid, without smell. When prepared according to the London Pharmacopœia it has a density of 1·063, and contains rather less than five per cent of potassa.

The protohydrate of potassa (the *lapis infernalis*, or potential cautery of the older authors) is obtained by evaporating the solution in a silver or iron capsule. When all the water, save the one atom attached to each one of potassa, is driven off, the hydrate is poured into moulds, where it solidifies.

Potassa, like the other alkalis, forms soaps with animal and vegetable oils. This subject will be recurred to in Books III. and IV.

Potassa, when free or combined, is recognised by its not giving a precipitate with hydrosulphurets and carbonates, and by tartaric and perchloric acids precipitating with it tartrates and perchlorates. Chloride of platinum also throws down a yellow precipitate with potassa or potassa salts.

3. Peroxide of Potassium.

This may be formed by burning potassium in an excess of oxygen. It is a yellow powder, which on meeting with water is immediately resolved into potassa and oxygen. It is of little importance.

4. Chloride of Potassium.

This compound is formed when potassium burns in chlorine, and also when hydrochloric acid acts upon potassa; the results in the latter being water and the compound in question. It is also the residue when chlorate of potassa is decomposed by heat (*see* p. 98.).

This chloride was formerly called the febrifuge salt of Sylvius, and in common language, even at present, is named the muriate of potassa. It has a saline and bitter taste, is insoluble in alcohol, but dissolves in three parts of water at 60°, and in less of boiling water. It is obtained as a residue in various chemical processes, and is used in the manufacture of alum.

5. IODIDE OF POTASSIUM.

Iodide of potassium is formed with evolution of heat and light when potassium is heated in contact with iodine. It is the sole residue also after decomposing iodate of potassa with heat, in a manner analogous to the decomposition of chlorate of potassa, noticed in the preceding paragraph. By neutralising potassa with hydriodic acid it may be obtained in solution. Another mode of procuring it, and the one ordered by the Pharmacopœia, is to digest iodine with iron filings in water, and to decompose the resulting iodide of iron with as much potassa as sufficient to precipitate the oxide of iron.

Iodide of potassium fuses when heated, and vaporises at a heat below redness. It is very soluble in water and alcohol. It has no smell, but has an acid saline taste. It is often much adulterated. It consists of one atom of iodine and one of potassium. It is extensively used in medicine.

6. Sulphurets of Potassium.

There are five sulphurets of potassium, any of which may be formed by heating the two elements together. More economical modes are, however, adopted.

The *ter*sulphuret may be prepared by fusing at a low heat one part of sulphur and two of carbonate of potassa. Carbonic acid gas is driven off with effervescence; a portion of the sulphur unites with some of the oxygen contained by the potassa, forming sulphuric acid; and this unites with the undecomposed potassa, giving rise to the sulphate of potassa; while the remaining sulphur and the potassium of the decomposed potassa, by their union, form the tersulphuret of potassium, or liver of sulphur, as it used to be called.

This is a liver-brown, deliquescent substance, occasionally used in therapeutics. If a solution of it be treated with an acid a curious set of changes occurs: the water is decomposed, and its hydrogen unites with some of the sulphur, and forms sulphuretted hydrogen, and the oxygen joins with the potassium, the result being potassa; the remaining sulphur then falls down in a state of very minute division, and of a cream colour, and constitutes the *lac sulphuris* of the apothecary.

II. Sodium.

Sodium or Natrium occurs, although always in combination, most abundantly in nature. It is the principal saline ingredient in the waters of the ocean, and occurs in many minerals, and as a constituent of animal and vegetable structures.

Its discovery was made by Davy a few days after that of potassium. He procured it from soda, by means of galvanism; but it is now obtained by a process exactly similar to that just mentioned under potassium.

Sodium has a metallic lustre very like that of silver. At common temperatures it is soft, at 200° it melts, and it vaporises at a red heat. Its specific gravity is ·972, and it therefore swims upon water. When thrown upon that fluid it is rapidly oxidated, but it does not inflame like potassium, unless the water be hot, or if the bit of metal be prevented from moving about.

Its symbol is Na, and its atomic weight 23.

The compounds of sodium described in this section are—

		Formula.
1. *Protoxide of sodium*	- Sodium 1 atom 23+Oxygen - 1 atom 8=31	$Na+O$.
2. *Chloride of sodium*	- Sodium 1 atom 23+Chlorine 1 atom 35=58	$Na+Cl$.
3. *Sulphuret of sodium*	- Sodium 1 atom 23+Sulphur - 1 atom 16=39	$Na+S$.

1. PROTOXIDE OF SODIUM.

This is commonly called soda, and is prepared from the carbonate of soda exactly in the same manner as potassa from carbonate of potassa.

In its properties soda resembles potassa; but it is a weaker alkali, and all its salts are decomposed by potassa.

Soda may be distinguished from the other alkaline bases, 1st, by yielding, with sulphuric acid, a salt easily known by its taste and form to be sulphate of soda or Glauber's salts; 2nd, by all its salts being soluble (none, therefore, ever being precipitated); and, 3rd, by its salts burning, when exposed to the blowpipe, with a yellow colour.

There is a peroxide of sodium, but it is of no importance.

2. Chloride of Sodium.

This is the common salt of the vernacular. As sea water contains, on an average, two and a half per cent of it, the quantity in this form of solution must be enormous. Salt or brine springs, too, are very common, especially in the transition and cretaceous formations. Indeed nearly all spring water contains more or less of it. It is very extensively used as a necessary article of food, living structures essentially containing it.

The salt consumed in this country is procured by evaporating the water of the brine springs. Of these that in Cheshire is perhaps the most important. The brine is pumped into evaporating pans, to which heat is applied. Various commercial kinds of salt are thereby obtained, which, as well as those obtained from sea water, will be noticed in Book IV.

Chloride of sodium may, if wished, be formed either by the direct union of its elements, or by decomposing hydrochloric acid by carbonate of soda; the result of the latter process being, of course, chloride of sodium and water.

Chemically pure chloride of sodium has a pleasant, pure saline taste. It is permanent in the air, but if, as usually is the case, it contains magnesian salts, it is deliquescent. At a red heat it fuses, and upon cooling it becomes a transparent brittle mass. It dissolves in two and a half times its weight of water at 60°; but its solubility is not increased at higher temperatures. It is anhydrous, but a quantity of water is usually inclosed between its crystals; and hence, if thrown upon the fire, it decrepitates, owing to this water being converted into vapour and splitting the crystals.

There is an iodide of sodium prepared in a manner analogous to the mode of formation of iodide of potassium.

3. SULPHURET OF SODIUM.

There is more than one sulphuret of sodium, but the protosulphuret only has been studied. It may be formed by the direct union of its constituents, and is chiefly interesting from its being a constituent of lapis lazuli. Artificial ultramarine may be prepared by heating this substance along with a mixture of silica and alumina.

* * *

III. LITHIUM.

This element is the metallic base of lithia, and is only present in a very few minerals. It is only a subject of interest to professed chemists.

* * *

IV. Barium.

This metal is only found (and then always combined) in the mineral kingdom, and never as a constituent of animal or vegetable structure. Even in the mineral world it is not very abundant.

It was separated from its oxide by Davy, who employed the galvanic battery for the purpose. It is a dark gray-coloured metal, much heavier than water. Like potassium and sodium, it has a strong affinity for water, and if exposed to the air or water, becomes rapidly oxidated.

Its symbol is Ba, and its atomic weight 68.

The compounds of barium described in this section are as follow:—

		Formula.
1. *Protoxide of barium* -	Barium 1 atom 68+Oxygen 1 atom 8=76	Ba+O.
2. *Sulphuret of barium* -	Barium 1 atom 68+Sulphur 1 atom 16=84	Ba+S.*

1. Protoxide of Barium, or Barytes.

This is so named from *βαρὺς heavy*, owing to the great weight of the minerals that contain it. It was discovered by Scheele in 1774. It is best prepared by heating the nitrate of barytes in a crucible, upon which the heat expels all the nitric acid.

Barytes is a grey powder, having a specific gravity of about 4. It has a caustic, alkaline taste; converts vegetable blues to greens, and neutralises the strongest acids. It is, however, much less caustic and much less soluble in water than potassa and soda.

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* Other binary compounds of barium with the non-metallic elements are the peroxide, the chloride, the iodide, &c.

Baryta has a remarkable affinity for water; and combines with that fluid with noise and the evolution of great heat, accompanied sometimes with incandescence; the result being the formation of a hydrate of barytes.

An aqueous solution of barytes is a good test. of the presence of carbonic acid in the air. The carbonic acid unites with it, and falls down to the bottom as a white, insoluble carbonate of barytes.

All the soluble salts of barytes are poisonous to man.

Barytes is distinguished from all other substances by, 1st, dissolving in water and forming an alkaline solution; and, 2nd, by all its soluble salts being precipitated as white carbonate of barytes by alkaline carbonates and as sulphate of barytes (which is insoluble in acid or alkaline solutions) by sulphuric acid or a soluble sulphate.

2. SULPHURET OF BARIUM.

This is prepared by heating heavy spar, which is sulphate of barytes, along with charcoal. The charcoal combines with the oxygen, forming carbonic acid, and sulphuret of barium is left behind. It is principally used as a source of the other preparations of barium. Thus, to obtain the chloride the sulphuret is dissolved in hydrochloric acid, the result being sulphuric acid and the chloride of barium.

V. Strontium.

At Strontian, in Argyleshire, a mineral occurs which was proved by the late Professor Hope to be the carbonate of a new earth, to which he gave the name of strontia. Subsequently this earth was found to be the protoxide of a metal, strontium.

Strontium and its compounds are exactly analogous with barium and its chemical combinations. Strontia is remarkable for communicating to flame a splendid crimson colour, and hence it is used by the makers of fireworks, and also for producing red signal lights.

Its symbol it Sr, and its atomic weight 44.

* * *

VI. Calcium.

The existence of this metal was established by Davy, who procured it from lime by a process similar to that in which barium is obtained. It is a white metal, sinks in water, and is rapidly oxidated by that fluid.

Its symbol is Ca, and its atomic weight 20.

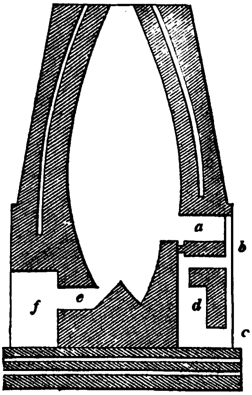
The compounds of calcium described in this section are as follow :—

		Formula.
1. *Protoxide of calcium* -	Calcium 1 atom 20+Oxygen - 1 atom 8=28	$Ca+O$.
2. *Peroxide of calcium* -	Calcium 1 atom 20+Oxygen - 2 atoms 16=36	$Ca+2O$.
3. *Chloride of calcium* -	Calcium 1 atom 20+Chlorine 1 atom 35=55	$Ca+Cl$.
4. *Fluoride of calcium* -	Calcium 1 atom 20+Fluorine 1 atom 18=39	$Ca+F$.
5. *Sulphuret of calcium.*	There are three sulphurets.	
6. *Phosphuret of calcium.*		

1. Protoxide of Calcium, or Lime.

This important earth is most widely distributed in nature. In the inorganic kingdom it is principally associated with sulphuric and carbonic acids, forming vast beds of gypsum, limestone, and chalk. It is also found combined with fluorine, phosphoric and arsenious acids. United with phosphoric acid it forms the bones of the higher animals, and with carbonic acid the shells of the lower. Indeed, the vast mountains and deposits of carbonate of lime, or limestones, are nothing but the aggregated shells of myriads of invertebrated animals that once lived in these localities.

Fig. 38.



Protoxide of calcium, lime or quick lime, is always obtained from the native carbonate. To obtain it perfectly pure, pieces of Carara marble should be placed in a crucible (loosely covered, that the carbonic acid may escape), and heat applied. The carbonate of lime is of course resolved into lime and carbonic acid, which latter passes away. On the large scale lime is obtained by burning it in kilns such as that shown in *fig.* 38. *a* is the fireplace upon which the fuel is placed, *b* the opening for the draught of air *c*, and *d* the ash-pits. The carbonate of lime is introduced at the top; and the lime, freed from the carbonic acid, is taken out at *e* and *f*. The flame only enters the kiln, and then no ashes from the fuel mix with the lime.

Lime is a brittle, white earth, having a specific gravity of 2·3. When heated to full redness it becomes powerfully phosphorescent. It is one of the most infusible bodies known, being fused with difficulty even by the heat of the oxy-hydrogen blowpipe. It has a very strong affinity for water, with which substance it forms a white hydrate, the combination being attended with so much heat as to char wood and inflame gunpowder. In this manner ships freighted with lime have been set on fire by water penetrating to the hold. This hydrate of lime is in common language called *slaked* lime.

Hydrate of lime consists of one atom of lime and one of water. It is very slightly soluble in water, and, a very singular fact, it is more soluble in cold water than in hot. Thus, one grain of lime requires for solution —

778 grains of water, at 60°,
972 do. do. 130°,
1270 do. do. 212°.

In consequence of this property, on heating a solution of lime water that has been made in the cold, lime is deposited.

The solution of hydrate of lime is called lime water. It has a harsh, acrid taste, and converts vegetable blue colours into green. It has a strong affinity for carbonic acid, and forms with it an insoluble carbonate. Hence it should be protected from the air. Carbonate of lime, however, is soluble in an excess of carbonic acid, and it is in this manner that a small quantity of carbonate of lime is held in solution by almost all spring water.

The most delicate test for salts of lime is oxalate of ammonia. On adding this substance to a solution of salt of lime, insoluble oxalate of lime is thrown down. Oxalic acid only forms an insoluble salt with lime, barytes and strontia. That it is the oxalate of lime is easily proved by dissolving it in nitric acid and evaporating. The nitrate of lime thus obtained is deliquescent, and soluble in alcohol, while the nitrates of barytes and strontia are neither.

Salts of lime, when heated before the blowpipe, communicate to the flame a dull brownish red colour.*

2. PEROXIDE OF CALCIUM.

This may be prepared by adding deutoxide of hydrogen to lime water. It is of no importance.

3. CHLORIDE OF CALCIUM.

This substance exists in sea water and many saline springs. It may readily be formed by dissolving carbonate of lime in hydrochloric acid. In this decomposition the hydrogen of the acid unites with the oxygen of the lime, forming water; the carbonic acid is given off, and the chlorine and calcium unite together and dissolve in the water. By evaporation of the fluid a hydrate of chloride of calcium is obtained. The water may be driven away from this by heat, but on exposure to the air it again rapidly absorbs water. Indeed, its principal use is owing to its affinity for water; and it is employed to remove water from liquids, as in the rectification of alcohol; to dry gases; and to form a frigorific mixture with snow.

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* Lime is of great importance in the arts as mortar, and as an application to the soil to increase its fertility.

Mortar consists of lime, sand, and hair. The sand prevents the irregular cracking of the lime as it dries, and the bits of hair serve to bind the whole mass together. The coherent property of the mortar depends upon the hydrate of lime gradually drying, so that the stones are retained together in the same manner that boards are by the drying of glue.

Hydraulic mortar used for subaqueous buildings is an entirely different substance from common mortar. It is a chemical compound of lime, silica, oxide of iron, and alumina.

The action of lime as a fertiliser is not intelligible until the soil and plants have come before our notice. Its consideration is deferred until Book IV.

4. Fluoride of Calcium.

This occurs native, and often accompanies ores of lead and tin. It is common in the mines of Derbyshire, and is often collected; and many of the specimens found in that locality are remarkable for their size and the beauty of their colours. Small quantities of it are present in all fertile soils; it is found in the structure of vegetables, and forms a small portion of the bones of animals. It is rather a curious fact that some of the fossil bones have been found to contain a larger proportion of it than modern bones do.

It may be artificially prepared in a variety of ways. It is employed as a flux in metallurgy, for which it is qualified, by forming with silicious and aluminous minerals fusible compounds, for making vases, and as the source to the chemist of fluorine.

5. Sulphuret of Calcium.

There is more than one sulphuret of calcium. The protosulphuret is formed by heating together sulphate of lime with charcoal. All the oxygen, both of the sulphuric acid and the lime, combines with the carbon of the charcoal, and flies off as carbonic acid. The sulphur and calcium that are left combine together, forming a white powder with a reddish tint. It becomes phosphorescent after exposure to light, and has sometimes been called Canton's phosphorus.

6. PHOSPHURET OF CALCIUM.

If vapour of phosphorus be passed over quicklime at a low red heat, a brown substance is formed that consists of a mixture of phosphate of lime and phosphuret of calcium. If these be put into water mutual decomposition takes place, and phosphate of lime and phosphuretted hydrogen are formed, the latter being in the form that spontaneously inflames.

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VII. MAGNESIUM.

Magnesium, not free, however, but always combined, occurs very abundantly in nature. It is a constituent of many minerals, is always present in soil, is present in sea water, and is a necessary ingredient in animal and vegetable structures. Its existence was demonstrated by Davy, who procured it by means of galvanism; but the best mode of obtaining it is that of Bussy, which consists in mixing together chloride of magnesium and potassium; upon which the chlorine leaves the magnesium for the
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potassium, and the result is chloride of potassium and metallic magnesium.

Magnesium is a brilliant white metal with a lustre resembling that of silver. It is very malleable. It fuses at a red heat. It is oxidised by exposure to damp air, but undergoes no change when kept in dry air. It may be boiled in water without being affected. If heated to redness in air or oxygen it inflames, and its oxide or magnesia is formed. Placed in dilute acids it dissolves, hydrogen is liberated, and a salt of its oxide is formed, the oxygen having been derived from the decomposed water.

It symbol is Mg, and its atomic weight 12.

The compounds of magnesium described in this section are as follows:—

		Formula.
1. *Oxide of magnesium* - -	Magnesium 1 atom 12+Oxygen - 1 atom 8=20	$Mg+O$.
2. *Chloride of magnesium* - -	Magnesium 1 atom 12+Chlorine 1 atom 36=48	$Mg+Cl$.
3. *Bromide of magnesium* - -	Magnesium 1 atom 12+Bromide 1 atom 78=90	$Mg+Br$.

1. Oxide of Magnesium, or Magnesia.

This, the only compound of oxygen and magnesium, is a protoxide. It is always obtained by subjecting the carbonate of magnesia (a compound hereafter to be described) to a full red heat, upon which the carbonic acid is expelled and pure magnesia left behind.

It is a light white powder, having neither taste nor smell. It is almost infusible, but, when strongly heated, it becomes powerfully phosphorescent. It is very insoluble in cold water, requiring more than 5000 times its weight of that fluid at 60° to dissolve it, and, like lime, it is still less soluble in warm water. The solution browns turmeric paper slightly, but it neutralises perfectly the strongest acids.

Magnesian salts are known by their alkalinity, the great solubility of the sulphate, and a solution containing it being precipitated by alkalies. The phosphate of ammonia and soda produces in a liquid containing magnesia a deposition of crystals of phosphate of ammonia and magnesia.

2. CHLORIDE OF MAGNESIUM.

This may be prepared by mixing hydrochloric acid with carbonate of magnesia. It is interesting only as the source of magnesium. It exists in sea water.

3. BROMIDE OF MAGNESIUM.

This, too, exists naturally in sea water, and is of importance only as the source of bromine. It may be formed artificially by mixing hydrobromic acid and magnesia, the result being water and the compound under notice. It has an extremely bitter taste.

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(B.) VIII. ALUMINUM.

This metal occurs very abundantly in nature, but always in a state of combination. It may be separated from its chloride in the same manner that magnesium is obtained from the chloride of that metal.

It is a grey powder, the scales of which have a
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metallic lustre. It does not decompose water at ordinary temperatures, and indeed but slowly at the boiling point; but it dissolves rapidly in diluted acids, hydrogen being given off, and the aluminum being oxidated.

Its symbol is Al, and its atomic weight 13.

The compounds of aluminum described in this section are as follows:—

				Formula.
1. *Sesquioxide of aluminum* -	Aluminum 2 atoms 26+	Oxygen 3 atoms 24=	50	$2Al+3O$.
2. *Sesquichloride of aluminum* -	Aluminum 2 atoms 26+	Chlorine 3 atoms 105=	131	$2Al+3Cl$.
3. *Sesquisulphuret of aluminum* -	Aluminum 2 atoms 26+	Sulphur 3 atoms 48=	74	$2Al+3S$.

Sesqui-salts, it may be premised, are salts the elements of which do not combine one atom of the one with one, two, three, four, &c. of the other; but apparently one atom of one combines with one and a half of the other. But a half-atom is an impossibility, an atom of matter being the last possible division of it. Each atom of these compounds, therefore, is believed to consist of three atoms of the one constituent and two of the other. The name applied to such compounds, *sesqui*, is objectionable, but custom has sanctioned it.

1. SESQUIOXIDE OF ALUMINUM, OR ALUMINA.

This is the only oxide of aluminum, and it exists in nature even more abundantly than lime. Indeed, with the exception of silica, it is the principal ingredient in almost all known rocks; and as soils are formed by the decomposition of superficial rocks, it is the principal ingredient in all soils, save a few thin ones formed over the limestones and the sands of deserts. The common name, as it exists in soils, is *clay.*

Both in rocks and soils it exists generally combined with silicic acid, and also, but much more rarely, with the sulphuric and phosphoric acids. Sometimes, indeed, it occurs pure, or at any rate nearly so, forming the ruby and the sapphire. It also occurs pretty extensively combined with sulphuric acid and potassa, constituting the alum of commerce.

It is from this last-mentioned compound that alumina is procured. To a solution of it carbonate of potassa is added; the sulphuric acid, combined with the alumina, immediately attaches itself to the potassa, carbonic acid being evolved and hydrate of alumina precipitated. This hydrate is a gelatinous mass, but by heating the water is expelled, and the alumina obtained pure. It occupies, however, much less space than the hydrate.

Alumina has neither taste nor smell, and is quite insoluble in water. It is difficult of fusion, but not so much so as lime. It has a remarkable affinity for water, attracting that fluid from substances that contain it. It is owing to this property that a piece of dry alumina, as a tobacco pipe for example, adheres so closely to the lips. It is apparently capable of acting both as an acid and an alkali, forming salts with acids, and also uniting with the alkaline bases as potassa, soda, &c. But in neither of such cases are the salts formed neutral with regard to test paper.

Alumina, mixed with water, becomes a soft tenacious mass, that may be moulded into shapes and then have the water expelled by the application of heat. This renders it the substance selected for making bricks, earthenware, and porcelain.

Another remarkable and very useful property that it possesses is its tendency to unite with almost all organic matters, and with the immediate results of the putrefactive decomposition of organic matters. Its action in the former of these respects is invaluable to the dyer; and, owing to its action in the latter, water containing organic matter or ammonia that have been formed upon the surface of the earth loses its impurities ere it can reach the springs.

Two circumstances distinguish alumina from all other substances. First, it is precipitated from acids as a hydrate by all the alkaline carbonates and by ammonia; and, secondly, pure potassa or soda precipitate it; but this precipitate is redissolved by an excess of the alkali.

2. Sesquichloride of Aluminum.

By dissolving alumina in hydrochloric acid a hydrate of sesquichloride of aluminum is obtained. By applying heat to this it is decomposed, and again separated into hydrochloric acid and alumina.

Anhydrous sesquichloride can only be obtained by igniting charcoal along with alumina, and passing a stream of chlorine over the mixture. The carbon of the charcoal unites with the oxygen of the alumina, and the aluminum and chlorine form a compound. This chloride is interesting only as the chemical source of aluminum.

3. SESQUISULPHURET OF ALUMINUM.

This may be formed by heating aluminum in the vapour of sulphur. The moment it comes into contact with water it is resolved into alumina and sulphuretted hydrogen, that is say, $(2Al+3S)$ with $+(3H+3O)$ give $(2Al+3O)$ and $(3H+S)$.

* * *

IX. GLUCINUM, YTTRIUM, THORIUM, ZIRCONIUM, CERIUM, LANTHANUM.

All these are metals that occur very rarely and in small quantities. Glucinum is found in the emerald and the beryl, yttrium in the Bohemian garnet, thorium in very scarce minerals, zirconium in the hyacinth, and cerium and lanthanum associated together in rare minerals.

* * *

X. MANGANESE.

Like all the metals previously enumerated, manganese is only a chemical curiosity, never occurs free in nature, and is put to no use. In a state of combination it is extensively diffused, and some of its compounds are employed in the arts.

Manganese is prepared with difficulty by deoxidating one of its oxides by means of charcoal. It is a hard brittle metal of a greyish colour. It has a strong affinity for oxygen, and tarnishes on exposure to air;

and slowly decomposes water at a red heat, or, if sulphuric acid be present in the water, it effects the decomposition rapidly. In the one case oxide of manganese and in the other sulphate of oxide of manganese are found, hydrogen being in both instances evolved.

Its symbol is Mn, and its atomic weight about 27.

The compounds of manganese described in this section are as follows:—

							Formula.
1. *Protoxide of manganese* -	Manganese	1 atom	27+Oxygen -	1 atom	8=	35	$Mn+O$.
2. *Sesquioxide of manganese* -	Manganese	2 atoms	54+Oxygen -	3 atoms	24=	78	$2Mn+3O$.
3. *Peroxide of manganese* -	Manganese	1 atom	27+Oxygen -	2 atoms	16=	43	$Mn+2O$.
4. *Manganic acid*	Manganese	1 atom	27+Oxygen -	3 atoms	24=	51	$Mn+3O$.
5. *Hypermanganic acid* -	Manganese	2 atoms	54+Oxygen -	7 atoms	56=	110	$2Mn+7O$.
6. *Protochloride of manganese*	Manganese	1 atom	27+Chlorine	1 atom	35=	62	$Mn+Cl$.
7. *Perchloride of manganese* -	Manganese	2 atoms	55+Chlorine	7 atoms	245=	300	$2Mn+7Cl$.

1. Protoxide of Manganese.

This may be prepared in a variety of manners. One mode consists in heating to redness some of the higher oxides, and passing a stream of hydrogen, which abstracts as much oxygen as to leave the protoxide behind. It is usually of a greenish grey colour. It has no action on vegetable colours; but, on the contrary, combines with acids and forms neutral salts, many of which are distinguished by a peculiar rose colour.

This same oxide may be obtained as a hydrate by precipitating one of its soluble salts by an alkali. This is at first of a white colour, but it rapidly absorbs oxygen from the air and becomes brown.

2. SESQUIOXIDE OF MANGANESE.

This occurs native, and may be artificially procured by exposing the peroxide to heat. Hence, it is this oxide that is left in the retort when oxygen is prepared from the peroxide by adding to it sulphuric acid. It is also the brown substance mentioned above, as being found when the hydrated protoxide is exposed to the air.

It can unite with acids and form salts, which are of a deep red colour, and which are also isomorphous with those of alumina.

This oxide stains glass purple or violet.

3. PEROXIDE OF MANGANESE.

This is often called the black oxide, and is the common form in which manganese occurs in nature. It may be artificially procured by exposing nitrate of oxide of manganese to a red heat, by which the nitric acid is expelled.

It is a black substance, that does not change by exposure to the air, is not soluble in water, and does not unite with either acids or alkalies; but, when heated with sulphuric acid, it is decomposed, and gives off half its oxygen, sulphate of the protoxide being left behind. Its use in obtaining chlorine is mentioned at p. 106.

It is employed in the arts to obtain chlorine, and also to afford oxygen to iron in glass making, as will be afterwards noticed.

4. MANGANIC ACID.

This substance, when combined with potassa, as it always is, is sometimes named the mineral chameleon, on account of the curious changes of colour it undergoes when dissolved in water. It is prepared by heating hydrate of potassa with peroxide of manganese. If the manganate thus obtained be put into water, a green solution is procured. This soon becomes blue, purple, and red, and at last a brown powder subsides to the bottom, leaving the liquid colourless.

The explanation of these changes of colours is this: manganate of potassa is green, and hence the green hue first observed. It is, however, very liable to decompose water and become the permanganate of potassa, which is red, and the blue and purple tints are formed by the mixture of this colour with the green. The brown powder is hydrated peroxide of manganese.

Manganic acid has never been obtained alone, as the moment it is placed in an uncombined state it is resolved into peroxide and oxygen.

5. HYPER- OR PER-MANGANIC ACID.

This, or, to speak correctly, the permanganate of potassa, is obtained by diluting a solution of manganate of potassa with hot water. Peroxide of manganese falls to the bottom, and red permanganate of potassa is formed.

Two other compounds of manganese and oxygen exist of a very complex nature, and of little importance. Indeed, the above five are enumerated more to illustrate the principle, that when a metal forms many oxides, those containing least of the oxygen act as bases, uniting with acids to form salts; while those that contain a greater quantity act as acids, than for any importance that most of them are of.

6. PROTOCHLORIDE OF MANGANESE.

This may be obtained by digesting the peroxide in hydrochloric acid, the result being the chloride of manganese and water. It is a pink-coloured substance, and very deliquescent.

7. PERCHLORIDE OF MANGANESE.

This appears to be formed when hydrochloric acid and peroxide of manganese are mixed together without heat. It is a substance of a green colour, and is, by the application of moderate heat, changed into protochloride and free chlorine.

This last-mentioned compound is introduced as an illustration of the general law not hitherto noticed, that the combination of metals with chlorine, &c., are generally in the same proportions as the compounds, or some of the compounds of the same metal with oxygen. The protochloride and perchloride of manganese are analogous to the peroxide and the permanganic acid.

The detection of manganese is easy. Ammonia throws down from a solution of its salts a white precipitate, which passes first to light and afterwards to dark brown, and which is the hydrated protoxide.

Sulphuretted hydrogen precipitates the flesh-coloured sulphuret of manganese.

* * *

(C.) XI. Iron, or Ferrum.

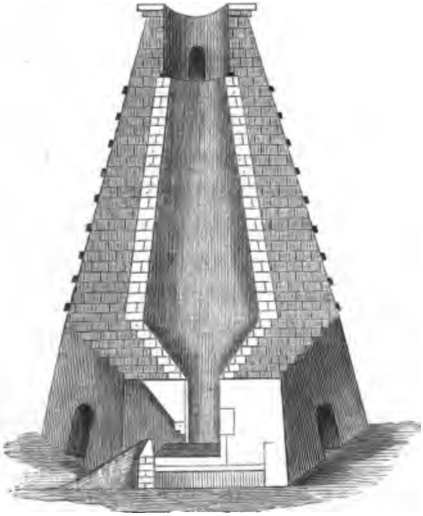
Iron, the Mars of the alchemists, has been known from the remotest antiquity. It is the most useful, as well as, fortunately, the most abundant of the metals. It is contained in animal and vegetable structures, is present in all fertile soils, and often in springs; it sometimes, but in minute quantity, occurs native; and its ores, of various kinds, are found in very large quantities. Of these, the most important are the red oxide or *hematite*, the hydrated peroxide or brown hematite, the black oxide or magnetic iron ore, and the carbonate, which, being generally mixed with alumina, silica, lime, &c., is called the clay-iron ore. Of these, the former are commonly found in primary formations; while the last occurs in secondary ones, particularly in that of the coal.

The native iron sometimes is found in small quantities in veins, but also in blocks lying upon the surface of the earth. These latter are not of terrestrial origin, but have fallen upon this planet; in all probability from volcanos in the moon. The iron in them is always associated with nickel.

Pure iron is very readily obtained from the oxides by simply heating them along with some carbonaceous matter, and the iron thus obtained is of superior quality. The reduction of the carbonate or clay-iron ore is a much more complex operation; and the difficulty is increased on account of different ores containing different amounts of silicious matter, &c.

The first step in the process of reducing clay-iron ore is to roast it, and thereby get rid of the carbonic acid and the water that it contains. The ore is then powdered and introduced by the top into the blast furnace, *fig.* 39. By means of a forcing pump, air, previously heated to about 700°, is driven into it. Before this practice of using heated air was discovered, coke was introduced along with the ore for

Fig. 39.



fuel; but now raw coal is substituted, and the saving to the iron masters is immense.

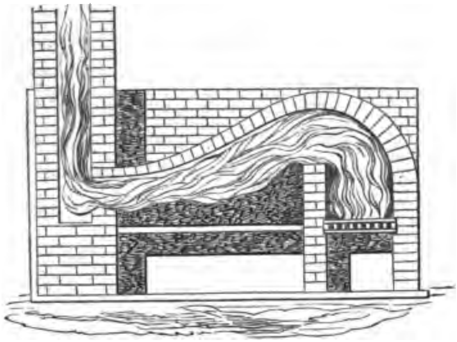
After the heat has been applied for some time the following changes take place: the silica, alumina, and lime unite together to form a fusible compound, technically called *slag*, and the oxide of iron parts with its oxygen to the fuel. The iron thus liberated falls, from its greater specific gravity, to the bottom; but in doing so it combines with a portion of the carbon of the fuel, and when it has reached the bottom it is a carburet. It is then drawn off, and constitutes the cast iron of commerce.

In order that the slag may be formed, and the process successful, it is necessary that the silica, lime, and alumina be in such proportions that they will form a soluble compound. This is never the case in the ore, and lime is always deficient. Hence, lime is invariably added to the mixture of ore and fuel. Some ores contain too little sand, and to such silicious matter is put. Very often different ores are mixed together, in order to obtain these three substances in their right proportions.

The carburet of iron, or cast iron, differs from pure iron in many important physical properties, and hence it is necessary to get rid of the carbon. This is done in a reverberatory furnace (see *fig.* 40.) by means of *puddling*. This consists in stirring the melted iron with a long iron rod, the object being to expose as much as possible of the iron to the air. The result of this is, that the carbon combines with the oxygen of the air to form carbonic oxide, which is driven up the chimney.

Cast iron is much more readily fusible than iron proper, and in consequence of this a curious phenomenon is witnessed during the puddling. The heat of the reverberatory furnace is quite sufficient to melt the cast iron, and the rod of the puddler moves easily enough to and fro in the molten mass. As the carbon is gradually expelled it becomes pasty, and the rod can hardly stir it, and at last it becomes quite solid. The iron is then rolled and beat.

Fig. 40.



For all economical purposes this iron is sufficiently pure. But it still contains about half per cent. of carbon, and also some silica. To obtain perfectly pure iron for chemical purposes, the best mode is to heat the pure oxide in a porcelain tube, and transmit a stream of hydrogen over it. The result of course is water and iron.

Of late it has been common to expedite the process of puddling by adding some substance to afford oxygen besides the air, as water or peroxide of manganese.

Iron has a grey colour, and a strong metallic lustre, that is capable of being much increased by polishing. Heated to redness it becomes soft and pliable, and may be incorporated or welded with another piece of red-hot iron by hammering. Its specific gravity is between 7 and 8. It is very infusible, only yielding to the highest temperature that may be produced by a wind furnace. It is attracted by the magnet, and may itself be rendered permanently magnetic; a property only possessed by one other metal, nickel. It can only maintain this magnetic property, however, within certain temperatures; and iron ceases to be attracted by the magnet at a red heat, and a steel magnet loses its polarity at the boiling point of almond oil.

The affinity of iron for oxygen is strong. It cannot, however, at a low temperature act upon dry air; but if moisture be present it steadily becomes oxidated or rusted. Heated to redness in the air it absorbs oxygen rapidly; but it combines with it in another proportion than when exposed at common temperatures to watery vapour.

Iron is inferior in durability and malleability to several metals; but in tenacity it exceeds them all.

Its atomic weight is 28, and its symbol Fe.

The compounds of iron described in this section are as follow:—

							Formula.
1. *Protoxide of iron* -	Iron -	1 atom	28+	Oxygen -	1 atom	8= 36	$\mathrm{Fe+O}$.
2. *Peroxide*[*] *of iron* -	Iron -	2 atoms	56+	Oxygen -	3 atoms	24= 80	$\mathrm{2Fe+3O}$.
3. *Black oxide of iron* -	Protoxide Sesquiox.	1 atom 1 atom	36 80	- -	- -	=116	$\mathrm{(Fe+O)+}$ $\mathrm{(2Fe+3O)}$.
4. *Protochloride of iron*	Iron -	1 atom	28+	Chlorine	1 atom	35= 63	$\mathrm{Fe+Cl}$.
5. *Sesquichloride of iron*	Iron -	2 atoms	56+	Chlorine	3 atoms	105=161	$\mathrm{2Fe+Cl}$.

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[*] Or sesquioxide.

						Formula.
6. *Protiodide of iron* -	Iron	- 1 atom	28+Iodine	- 1 atom	126=154	$Fe+I$.
7. *Sesqui-iodide of iron*	Iron	- 2 atoms	56+Iodine	- 3 atoms	378=434	$2Fe+3I$.
8. *Protosulphuret of iron*	Iron	- 1 atom	28+Sulphur	- 1 atom	16= 44	$Fe+S$.
9. *Sesquisulphuret of iron*	Iron	- 2 atoms	56+Sulphur	- 3 atoms	48=104	$2Fe+3S$.
10. *Bisulphuret of iron*	Iron	- 1 atom	28+Sulphur	- 2 atoms	32= 60	$Fe+2S$.
11. *Magnetic pyrites* -	Composition mixed.					
12. *Carburets* -	Composition undetermined.					

1. PROTOXIDE OF IRON.

This oxide is the base of carbonate of iron and of the sulphate or green vitriol. It may be obtained as a white hydrate by adding pure alkalis to some of its salts; but it cannot be procured pure in a dry state, on account of its great affinity for oxygen, which causes it to be converted into a higher oxide. If we attempt to form it by a process analogous to that employed for getting protoxide of manganese, the iron is reduced either to the black oxide or to the metallic state.

The protoxide combines with acids and forms salts, all of which are of a green colour. Instead of calling them sulphate, or nitrate, &c. of protoxide of iron, it is common to say the protosulphate, the protonitrate, &c., of iron. All the protosalts of iron are very prone to absorb oxygen and pass into the state of per-salts, from which they may be distinguished by their forming the white precipitate with alkalis.

2. SESQUIOXIDE OF IRON.

This is the crocus Martis of the older chemists, and is still often called the red oxide of iron. It is found native, constituting (combined with water) iron glances, hematite, ochre, &c. It may be artificially prepared in a variety of ways. The one employed by the London pharmacopeia consists in mixing together, in solution, carbonate of soda and sulphate of protoxide of iron. The sulphuric acid attacks the soda, the carbonic acid is given off, and the protoxide of iron combines with some oxygen of the air and becomes the hydrated sesquioxide. It should be observed that there is no carbonate of the peroxide.

Iron rust is hydrated sesquioxide.

Sesquioxide of iron is not attracted by the magnet. Almost all its salts are red. Its presence may be very decidedly detected. Alkalis (and, for the reason above stated, alkaline carbonates) precipitate it as a red hydrate. Infusion of nutgalls form with it the black tannate of iron or ink; with ferrocyanuret of potassium it makes Prussian blue, and the sulphucyanuret of potassium strikes with it a deep blood-red colour.

A few drops of nitric acid oxidate the protoxide and convert it into peroxide. Hence, it is usual in testing for iron to add a few drops of this acid, and apply the test for the sesquioxide.

3. Black Oxide of Iron.

This is the Ethiops martialis of the older writers, and is now often called the magnetic oxide. It is formed by heating iron at a high temperature; and the scales of a blacksmith's forge consist of it. It is attracted by the magnet, and is often itself magnetic.

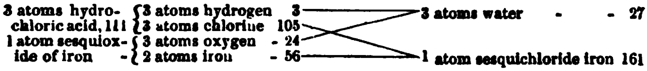
It is merely a mixture of the two other oxides; generally, perhaps, in the proportion of two parts of protoxide to one of sesquioxide. It may be artificially prepared by precipitating the solution of the two oxides in sulphuric acid.

4. Protochloride of Iron.

If iron be dissolved in hydrochloric acid, hydrogen is evolved, and this substance formed. By evaporation it may be obtained in crystals of a green colour.

5. Sesquichloride of Iron.

This is prepared by dissolving sesqiuoxide of iron in hydrochloric acid, the result being the compound under notice and water. Thus:—



This compound is very soluble in alcohol, and its tincture is pretty often employed in medicine.

6. Protiodide of Iron.

If water, iron wire, and iodine are mixed together, the two latter unite and form a green solution. It is necessary that the iron be in excess. By evaporation the iodide may be obtained in the form of an opaque iron-grey mass, having a faint metallic lustre and a styptic taste. It is very deliquescent, and dissolves readily in both water and alcohol. Its aqueous solution is very prone to take oxygen from the air, and to be resolved into sesqui-iodide and sesquioxide; and in order to prevent this a piece of iron wire must be kept permanently in it.

It is interesting mainly from its therapeutical properties, and the medicinal effects of iron and iodine which are supposed to be united in it.

7. Sesqui-iodide of Iron.

This may be obtained by exposing a solution of the iodide to the air, or by mixing in water iodine and iron; the former in excess. It is a volatile red compound, deliquescent, and soluble in water and alcohol.

Starch is turned blue by it, and hence it has been conjectured that it is more a mechanical admixture of iodine and protiodide than a chemical compound.

8. Protosulphuret of Iron.

Sulphur and iron have a strong affinity for one another. If an iron wire be heated to whiteness and touched with a stick of sulphur, the two combine; and if the iron be in excess the protosulphuret is formed. It is a black substance with a yellowish fracture. If dissolved in a dilute acid sulphuret of hydrogen is evolved, and a salt of the protoxide formed.

When moist, it has a very strong affinity for oxygen, which it absorbs, forming the protosulphate of iron. It sometimes occurs native in coal mines, and then occasionally absorbs the oxygen so fast, and with such intensity of chemical action, as to set the coal on fire.

9. SESQUISULPHURET OF IRON.

This corresponds to the sesquioxide. It may be formed by adding perchloride of iron to sulphuret of potassium, when mutual decomposition takes place. It is not attracted by the magnet.

10. BISULPHURET OF IRON.

In general when a substance combines in several proportions with oxygen, it combines in the same proportions with sulphur, and *vice versâ*. This compound is an exception to the latter part of this rule, there being no binoxide of iron. Bisulphuret of iron, however, occurs very abundantly, constituting the iron pyrites of mineralogists. It may be prepared artificially by heating together protosulphuret and sulphur. It is not magnetic.

11. MAGNETIC SULPHURET OF IRON.

There is more than one kind of a mixed sulphuret that is magnetic; but the most remarkable is that which corresponds to the magnetic oxide. It occurs native, and may be artificially formed by heating the bisulphuret to redness, and thus getting rid of a portion of the sulphur.

12. CARBURETS OF IRON.

There are three compounds of carbon and iron that are very distinct from one another—plumbago, steel, and cast iron.

Plumbago, or blacklead, as it is very improperly called, is found native, chiefly in primitive rocks and coal formations. It may be artificially produced by exposing iron with an excess of charcoal to violent heat. It has an iron-grey colour, a granular texture, and is soft and unctuous to the touch. Pure specimens contain about five per cent. of iron. Plumbago is used for making pencils and for polishing iron grates.

Steel contains about one and a half per cent. of carbon. It is prepared by surrounding bars of iron with charcoal, and exposing the two to a full red heat for many hours. It is less ductile and malleable than iron, but far harder and more elastic.

Cast iron is another compound of carbon and iron, containing about five per cent. of the former. It is easily fusible. When it becomes solid, after having been fused, it tends to crystallise, and in doing so expands powerfully. Hence, when poured into moulds in a liquid state, it fills up the minutest crevices; and it is extensively employed, in virtue of this property, to make castings, from which circumstance it derives its name.

If cast iron remain for a considerable time under water it becomes oxidised; the black oxide being the one that is formed, and the carbon remains behind under the form of a spongy mass, retaining very perfectly the figure of the original mass.

* * *

XII. NICKEL.

In the German mines an ore is found resembling, in external appearances, an ore of copper, but which resisted all the attempts of the miners to extract that metal from it. Hence, they termed it *kupper nickel, cheating copper.* It was subsequently found to possess a distinctive metal, which has been named nickel. It is chiefly interesting as being employed, along with zinc and copper, to form German silver.

A small quantity of nickel is always in the meteoric stones.

Nickel is of a white colour, approaching to that of silver, and of a very brilliant metallic lustre. It is ductile and malleable. It is attracted by the magnet, and can be rendered magnetic. It unites with oxygen in two proportions; but the protoxide only forms salts, all of which are of a green colour.

Nickel is readily recognised by yielding no precipitate with sulphuretted hydrogen, but by giving its hydrated green oxide by the pure alkalis, and by the same being precipitated by ammonia, but being redissolved in an excess of it, forming an azure solution.

The atomic weight of nickel is 29, and its symbol Ni.

* * *

XIII. COBALT.

The origin of the name of this metal is even more singular than that of the last. Kobald is the German name of the demon of the mines, and as the miners could not reduce the ores of cobalt they considered them under the protection of Kobald, and named them accordingly.

Cobalt exists in nature combined with arsenic and sulphur, and is universally associated with nickel. It is one of the constituents of meteoric stones. It is a brittle, reddish-grey metal. It forms two oxides; but only the protoxide forms salts, most of which are of a pink hue. The protoxide itself, when uncombined with water, has a light ash-grey colour, but in the state of hydrate a beautiful blue. Mixed with glass it communicates to that substance a bright blue. With alumina it forms the pigment cobalt blue, used as a substitute for ultramarine. Its combination with glass ground to a fine powder constitutes smalt, used for communicating a blue tint to writing paper, linen, &c.

Cobalt is very readily recognised by giving a blue tint to borax when melted with it.

There is a chloride of cobalt remarkable for writing with, its solution being white until dried before the fire and then becoming grass green. It is used for making sympathetic ink.

The atomic weight of cobalt is 29, and its symbol Co.

* * *

XIV. Zinc.

This metal was known to Paracelsus; but one of its ores was used in antiquity along with copper to form the alloy brass.

Zinc only occurs in the mineral kingdom. Its more important ores are calamine (the carbonate), blende (a sulphuret), and electric calamine (the silicate). When it is desired to extract it from the sulphuret, that ore is roasted, to expel the sulphur; during which process the zinc becomes oxidated. It is then mixed with carbonaceous matter and introduced into a crucible, the bottom of which is perforated by an iron pipe that terminates in a vessel of water placed in the room below. The zinc is reduced, and vaporised. It then falls down the tube and is condensed in the water. This is an instance of *distillation by descent.*

It is a bluish white metal of considerable brilliancy. Exposed to the air its surface oxidates, but the crust of oxide thus formed preserves the remainder of the metal in its metallic state. At low and high degrees of heat it is brittle; but at a temperature between 210° and 300° it becomes both malleable and ductile, and is readily rolled into thin leaves, or sheet zinc as it is called. This sheet zinc being cheaper than lead or tin is now extensively substituted for them. At a white heat zinc takes fire, burns with great brilliancy and with such violence, that the oxide thereby formed is mechanically carried into the air. Zinc is also oxidated by diluted sulphuric acid, the hydrogen of the water being evolved, as stated under hydrogen.

The atomic weight is 32, and its symbol Zn.

The compounds of zinc described in this section are as follow :—

		Formula.
1. *Protoxide of zinc*	- Zinc 1 atom 32 + Oxygen - 1 atom 8 = 40	$Zn+O$.
2. *Chloride of zinc*	- Zinc 1 atom 32 + Chlorine - 1 atom 35 = 67	$Zn+Cl$.
3. *Sulphuret of zinc*	- Zinc 1 atom 32 + Sulphur - 1 atom 16 = 48	$Zn+S$.

1. Protoxide of Zinc.

This has received the fantastical name of flowers of zinc, philosophers' wool, &c. It is formed during the solution of zinc in sulphuric acid, or it may be obtained by heating the carbonate to redness or by burning zinc. It is at ordinary temperatures a white powder; but when heated to redness it becomes yellow. It forms two classes of salts; one in which it is the base, and the other with alkalis, in which it acts the part of an acid. The members of the latter, however, are of no importance.

The presence of zinc in basic salts is known by this oxide being precipitated from its solutions as a white hydrate; by pure alkalis, which, however, redissolve it, if added in excess; and by hydrosulphuret of ammonia, causing a white precipitate, the hydrated sulphuret of zinc.

2. Chloride of Zinc.

This may be prepared by introducing zinc filings into chlorine gas; but is more conveniently procured by dissolving zinc in hydrochloric acid and evaporating the solution. It is a colourless salt, which at common temperatures is soft, and was in consequence of this property formerly called the butter of zinc; and it fuses at a heat a little above 212°.

Chloride of zinc is a powerful caustic, and its solution is one of the most effectual antiseptics and disinfectants that we have. It would appear to protect wood saturated with it from the dry rot; and it decomposes hydrosulphuret and carbonate of ammonia, and is hence usefully employed in hospitals, crowded ships, &c.

3. SULPHURET OF ZINC.

This occurs native, but may be artificially prepared by heating sulphur and zinc together. It is very variable in its colour, being sometimes black, or red, or yellow.

XV. CADMIUM.

This metal occurs very sparingly, and is found but in very small quantities in most ores of zinc. It is obtained from such by taking advantage of its greater volatility. It is a white metal, forms only one oxide, which combines with acids, forming white salts very much resembling those of the oxide of zinc. Cadmium salts, however, may be distinguished from zinc salts by giving with sulphuretted hydrogen a yellow (not *white*) precipitate, and by ammonia giving a white precipitate insoluble in an excess.

XVI. TIN, OR STANNUM.

Tin was perhaps the metal first known. Before mankind were able to work iron, their implements and weapons were made of an alloy of tin and copper; and by examining such in tumuli, &c., archæologists have been able to discriminate pretty accurately a pre-historic age of bronze.

Tin is the Jupiter of the alchemists.

The most common ore of tin is the native peroxide. The metal is obtained from this simply by heating it with carbonaceous matter.

Tin is the whitest of all the metals except silver. In its massive form, however, it has a yellowish hue. When rubbed it has a peculiar odour. It is very fusible, melting at 442°, and at a white heat it volatilises. When in contact with air or water it oxidates very slowly, and hence cooking utensils of copper are protected by a lining of tin. It is malleable, but very sparingly ductile.

Its atomic weight is 58, and its symbol Sn.

The compounds of tin described in this section are as follows: —

					Formula.
1. *Protoxide of tin*	- Tin 1 atom	58+Oxygen	- 1 atom	8= 66	$Sn+O$.
2. *Sesquioxide of tin*	- Tin 2 atoms	116+Oxygen	- 3 atoms	24=140	$2Sn+3O$.
3. *Peroxide of tin*	- Tin 1 atom	58+Oxygen	- 2 atoms	16= 74	$Sn+2O$.
4. *Protochloride of tin*	- Tin 1 atom	58+Chlorine	1 atom	35= 93	$Sn+Cl$.
5. *Perchloride of tin*	- Tin 1 atom	58+Chlorine	2 atoms	70=128	$Su+2Cl$.
6. *Sulphurets of tin*	- There are 3 sulphurets.				

1. PROTOXIDE OF TIN.

When chloride of tin in solution is mixed with an alkaline carbonate, carbonic acid is given off, and hydrochloric acid and hydrate of the protoxide of tin are formed; the latter of which falls to the bottom. The water may be driven off from it by drying at a temperature under 196°, and with the least possible amount of exposure to the air.

The most remarkable property of this oxide is its tendency to combine with oxygen and to become the peroxide. Heated in the open air it does so at once, and if when cold it is touched with a red-hot wire it absorbs the oxygen so fast as to inflame.

In virtue of this property it can throw down platinum, silver, and mercury from the solution of their salts. Its action in this respect upon peroxide of gold is singular; it takes away a portion of its oxygen, converting it into protoxide; and a powder, which is a mixture of peroxide of tin and protoxide of gold, falls down. This is of a purple colour, has long been known as the purple of Cassius, and when fused into glass or porcelain produces a superb purple-red colour.

Salts of the protoxide of tin (it is a basic oxide) are known from all other substances by giving this purple precipitate with gold, and by *not* being soluble in a solution of ammonia.

2. SESQUIOXIDE OF TIN.

This is a grey powder that form salts with acids; which salts, however, have been little examined. Its solution in hydrochloric acid strikes with gold the purple of Cassius, and it is soluble in solution of ammonia.

3. PEROXIDE OF TIN.

This oxide is formed whenever tin is allowed to combine freely with oxygen, and it exists largely in nature. It is very little disposed to unite with acids, and indeed, if dissolved by them, generally separates from them as a gelatinous hydrate. With the pure alkalis it acts as an acid, and forms soluble compounds called stannates.

It is formed artificially in two ways. First, by acting upon tin by diluted nitric acid. When this is done, nitrous acid is evolved, and hydrated peroxide of tin falls to the bottom. By heating this to redness, the water is expelled and the yellow peroxide is left behind. Secondly, by decomposing the solution of the perchloride by an alkali it is precipitated as a hydrate, the action that takes place being analogous to that which occurs in obtaining the protoxide from the protochloride.

But the peroxides obtained in these two manners differ very much in their properties, and are therefore now regarded as isomeric bodies. The first one is called α peroxide, and the second β peroxide.

The α kind is not soluble in either nitric or sulphuric acids, whether strong or diluted. Neither does it dissolve in hydrochloric acid; but this reagent changes it into an insoluble basic salt.

The β kind is quite soluble in hydrochloric acid, and in dilute nitric and sulphuric acids.

If the α kind be distilled with hydrochloric acid, it is converted into the β kind. On the other hand, the β kind, by boiling with nitric acid, becomes the α kind.

The hydrated peroxide reddens litmus and combines with alkalis and the β kind of it; The β kind only combines with acids. It is used in the arts, and is known as *putty powder;* and is likewise employed to communicate the requisite whiteness to the enamel of watch faces, &c.*

4. PROTOCHLORIDE OF TIN.

This compound may be procured by transmitting vapour of hydrochloric acid over metallic tin; hydrogen is evolved, and the protochloride obtained in solution, from which it may be separated in a solid form by evaporation. Protochloride of tin is remarkable for its affinity for both oxygen and chlorine, owing to its strong tendency to be resolved into bichloride or oxide. Owing to this property it reduces salts of mercury, gold, and silver to their metallic state, and the salts of copper, iron, and manganese to their lowest points of oxidation. It has a similar effect upon many organic compounds, the result of which is to deprive them of their colour; and hence it becomes of importance as a dyeing agent.

5. PERCHLORIDE OF TIN.

If bichloride of mercury and metallic tin be mixed together, and moderate heat applied, a colourless liquor distils over, which is the per- or bichloride of tin, or *fuming liquor of Libanius*, and metallic mercury is left in the retort.

This is employed in dyeing, and is commonly called the permuriate of tin. For commercial purposes it is prepared by dissolving tin in nitro-muriatic acid. Bichloride of tin possesses an influence over organic colours similar to that of alumina.

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* The peroxide of tin may be distinguished by the action of hydrosulphuret of ammonia, which produces a golden-yellow precipitate.

6. Sulphurets of Tin.

There are three sulphurets of tin, corresponding to the three oxides; but of these the most interesting is the bisulphuret. It is particularly so to the student of chemical history, as it was one of the substances that kept alive the hopes of the alchemists.

It is best prepared by a rather elaborate process. Tin, mercury, sulphur, and sal-ammoniac are rubbed up together in a mortar, and placed in a flask, which is heated in a sand-bath until no more fetid white vapours are given off. The heat is then raised to redness; upon which sulphuret of mercury and chloride of tin sublime, leaving the bisulphuret of tin behind.

The use of the mercury in this mixture is not very evident, but it seems to promote the combination of the tin and the sulphur. The sal ammoniac, by its evaporation, prevents the temperature becoming so high as to decompose the bisulphuret.

The bisulphuret obtained in this manner is in crystalline scales of a golden yellow colour and metallic lustre. The common name is *mosaic gold*.

* * *

XVII. Chromium.

This is an interesting metal, owing to the beautiful pigments and dyes formed by its compounds. It was discovered in 1797 by Vanquelin. It occurs in some minerals; but most abundantly in the chromate of iron, which consists of a mixture of the oxides of iron and chromium. The reduction of this oxide to the metallic state is difficult, and need not be described here. (See Chromate of Potassa.)

Chromium has an atomic weight of 28, and its symbol is Cr.

The compounds of chromium described in this section are as follow:

								Formula.
1. *Sesquioxide of chromium*	Chromium	2 atoms	56+	Oxygen	3 atoms	24=	80	$2Cr+3O$.
2. *Chromic acid*	Chromium	1 atom	28+	Oxygen	3 atoms	24=	52	$Cr+3O$.
3. *Sesquichloride of chromium*	Chromium	2 atoms	56+	Chlorine	3 atoms	105=	161	$2Cr+3Cl$.
4. *Oxychloride of chromium*	(See p. 214.)			-		.	=213	$(Cr+3Cl)+(2Cr+3O)$.

1. SESQUIOXIDE OF CHROMIUM.

This compound may be obtained by heating chromate of mercury to redness; upon which the oxide of mercury and half of the oxygen of the chromic acid are expelled, and the green sesquioxide left behind.

Sesquioxide of chromium is insoluble in water. Fused with borax or vitreous substances, it communicates to them a beautiful green colour; and this property of it is rendered available in the arts. The colour of the emerald, it may be mentioned, is owing to the presence of this oxide. It unites with acids to form salts, all of which are of a green colour.

Compounds of oxide of chromium may be recognised by giving a green precipitate with alkalis.

2. Chromic Acid.

This may be prepared by mixing a solution of bichromate of potassa with sulphuric acid. The acid and the potassa combine, and the chromic acid separates in the form of beautiful red crystals. It combines with bases and forms salts important on account of the colours that they possess.

Chromic acid has a tendency to very readily dispose of a portion of its oxygen, and thus become converted into the sesquioxide. If boiled with starch, sugar, and other organic compounds, it is thus converted; and by virtue of this property it destroys the colour of indigo and of many other vegetable and animal colouring principles, and is, in consequence, employed in calico printing.

Chromic acid is known by its striking a rich yellow pigment, the chromate of lead, when mixed with a soluble salt of lead.

3. Sesquichloride of Chromium.

When oxide of chromium is mixed with lamp-black and exposed at a red heat to a current of chlorine, carbonic acid and chloride of chromium are given off. The latter, in the form of very beautiful peach-blossom-coloured scales, is obtained, being sublimed in the cold part of the tube.

4. Oxychloride of Chromium.

This, although not a binary compound, is mentioned here. It is formed by melting together in a crucible chloride of sodium and bichromate of potassa. The melted mass is then poured upon a slab and broken to pieces, and the pieces are introduced into a retort and sulphuric acid poured upon them. Instantly sulphates of potassa and soda are formed, and the curious compound under our notice distils over as a red liquid.

XVIII. VANADIUM.

This metal is obtained from a very rare mineral, in which it is combined with lead. It is also found (but in very minute quantities) in some ores of iron.

(D.) XIX. TUNGSTEN, OR WOLFRAM.

This also is a very rare and unimportant metal. It forms, with oxygen and sodium, a curious compound, that very strikingly resembles gold.

XX. MOLYBDENUM, OSMIUM, TITANIUM, COLUMBIUM.

These are metals of rare occurrence, and of little interest.

XXI. ANTIMONY, OR STIBIUM.

The sulphuret of antimony was known to the ancients, and was applied by the Asiatic and Grecian ladies as a pigment to the eyebrows. It was reduced to the metallic state and its preparations administered to man by Basil Valentine. This experimenter poisoned some monks with it, and hence its name antimony (*anti-moine*) is said to be derived. Its proper chemical name, however, is stibium.

Antimony occurs but rarely native. Its only frequent ore is the sulphuret of stibium. The metal may be obtained by fusing this along with black flux.

It is a brittle metal of a whitish-blue colour, and possessing considerable lustre. It melts at about 800°, and at a very high temperature volatilises. Heated to a white heat in a close crucible, and then suddenly exposed to the air, it takes fire. During the combustion a white vapour rises, which condenses on cool bodies in the form of white crystals. These are sesquioxide of antimony, and are called in the old books *argentine flowers of antimony.*

The atomic weight of antimony is 64, and its symbol Sb.

The compounds of antimony described in this section are as follow: —

		Formulæ.
1. *Hydruret of antimony* - -	Antimony 1 atom 64+Hydrogen 3 atoms 3= 67	$Sb+3H$.
2. *Sesquioxide of antimony* -	Antimony 2 atoms 128+Oxygen - 3 atoms 24=152	$2Sb+3O$.
3. *Antimonious acid*	Antimony 2 atoms 128+Oxygen - 4 atoms 32=160	$2Sb+4O$.
4. *Antimonic acid* -	Antimony 2 atoms 128+Oxygen - 5 atoms 40=168	$2Sb+5O$.
5. *Sesquichloride of antimony* -	Antimony 2 atoms 128+Chlorine - 3 atoms 105=233	$2Sb+3Cl$.
6. *Sesquisulphuret of antimony* -	Antimony 2 atoms 128+Sulphur - 3 atoms 48=176	$2Sb+3S$.
7. *Oxysulphuret of antimony* -	A mixture of sesquioxide and sesquisulphuret.	

1. Hydruret of Antimony.

It is best prepared by mixing together water, tartar emetic (containing tartrate of antimony), sulphuric acid, and zinc. The water is decomposed as mentioned under hydrogen, and the liberated hydrogen combines with the antimony, and forms the compound under consideration.

It is a colourless, insoluble gas, which precipitates most metals from their solutions, but not copper. When burnt it is resolved into water and metallic antimony. This compound is of no importance in itself; but it is liable to be confounded with the hydruret of arsenic, which will come under our notice immediately.

2. Sesquioxide of Antimony.

This may be prepared by adding to an acid and a boiling solution of chloride of antimony carbonate of soda. As carbonic acid does not combine with antimony, this oxide is precipitated. It is a dirty-white powder, which, if heated, becomes yellow. If heated in the open air it acquires oxygen and becomes antimonious acid.

It combines with acids to form salts that have little stability. It is, however, one of the bases of an important compound salt, the tartrate of antimony and potassa, or tartar emetic.

Antimonial salts are known by the bright, orange-coloured precipitate of sulphuret of antimony produced by sulphuretted hydrogen.

3. Antimonious Oxide.

This may be formed by heating the sesquioxide in the open air. It is a white powder, which combines with alkalis, forming salts known as *antimonites.* It is not acted upon by the fire.

4. Antimonic Acid.

This is prepared by digesting antimony in nitric acid. It is a yellow powder, that reddens litmus paper and forms with alkalies salts called *antimoniates.* Heated in the fire an atom of its oxygen is expelled and antimonious acid is left behind.

5. Sesquichloride of Antimony.

This compound has been long known, and was formerly called the butter of antimony. It may be formed by throwing antimony in powder into a jar of chlorine; combustion ensues, and the sesquichloride is generated. At common temperatures it has the consistence of butter. It is a caustic.

6. Sesquisulphuret of Antimony.

This is the most abundant ore of antimony. It may be artificially generated by transmitting sulphuretted hydrogen through any solution of an antimonial preparation. Artificially prepared, sulphuret of antimony has an orange colour.

It may be mentioned that the oxides of antimony and their compounds are all soluble in an excess of hydrochloric and tartaric acids.

7. OXYSULPHURET OF ANTIMONY.

If the sesquisulphuret of antimony be boiled with an alkaline liquor, a red powder (the oxysulphuret of antimony, or kermes mineral) is precipitated. The reaction that takes place consists in a portion of the oxygen of the potassa (or whatever alkali be employed) passing to the antimony and a portion of the sulphur to the potassium. If sulphuric acid be added to the remaining liquid, golden sulphuret of antimony is precipitated. This compound contains more sulphur than the kermes mineral. If the sulphuric acid be put to the liquor before the kermes has had time to separate, the precipitated sulphuret of antimony, as it is called, and which is a mixture of the two preceding, is thrown down.

XXII. ARSENIC.

Metallic arsenic sometimes occurs native, but is usually associated with iron or cobalt. The arsenical ore is roasted in a reverberatory furnace, and the arsenic, from its volatility, is expelled; combines with oxygen as it rises, and condenses on the sides of the chimney. This oxide is mixed with black flux and heated in a crucible; over which is luted an empty crucible into which the metallic arsenic passes, the oxygen of the oxide having combined with the carbon of the flux.

Arsenic is an extremely brittle metal of a strong lustre and a grey colour. Its specific gravity is nearly 6. Heated up to 350° it sublimes without previously liquefying, and at this time it emits a garlic odour; a property which no other metal possesses except perhaps zinc. In close vessels it may be sublimed without change, but if exposed to the air it becomes oxidated. It tarnishes by exposure to the air, and becomes converted into a grey powder, which is a mixture of metal and protoxide.

Its atomic weight is 38*, and its symbol As.

The compounds of arsenic described in this section are as follow: —

							Formula.
1. *Hydruret of arsenic*† -	Arsenic	2 atoms	76+Hydrogen	3 atoms	3=	79	$2As+3H$.
2. *Sesquioxide of arsenic*‡ -	Arsenic	2 atoms	76+Oxygen -	3 atoms	24=	100	$2As+3O$.
3. *Arsenic acid* -	Arsenic	2 atoms	76+Oxygen -	5 atoms	40=	116	$2As+5O$.
4. *Sesquichloride of arsenic* -	Arsenic	2 atoms	76+Chlorine	3 atoms	105=	181	$2As+3Cl$.
5. *Protosulphuret of arsenic* -	Arsenic	1 atom	38+Sulphur	1 atom	16=	54	$As+S$.
6. *Sesquisulphuret of arsenic* -	Arsenic	2 atoms	76+Sulphur -	3 atoms	48=	124	$2As+3S$.

1. HYDRURET OF ARSENIC, OR ARSENIURETTED HYDROGEN.

This gas is exceedingly poisonous, has proved fatal to several experimenters, and must be prepared and examined with the greatest caution. In *fig.* 41. *a b* is a bent glass tube fastened to a support. A glass jet *c* is fixed into the shorter tube. This jet is furnished with a cock. At the bend *d* is placed a piece of zinc. To a solution of arsenious acid sulphuric acid is added, and this mixture is poured into the long leg. The water is decomposed and the hydrogen and arsenic unite and form a gas which is driven into the short leg; and from this it is driven whenever the cock is turned.

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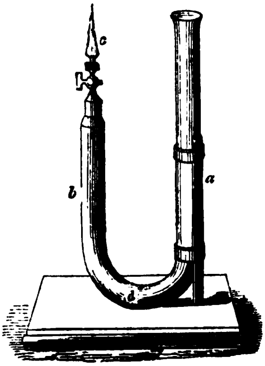
* Sometimes stated as 75.

† Arseniuretted hydrogen.

‡ Arsenious acid, or white arsenic.

Fig. 41.



Arseniuretted hydrogen is colourless, extinguishes burning substances, but is itself kindled by them, burning with a blue flame. The result of this combustion is water and metallic arsenic; the latter of which may be obtained by holding a piece of cold glass, &c., in the flame.

2. Arsenious Acid, or Sesquioxide of Arsenic.

This oxide of arsenic is remarkable among low metallic oxides as having no basic but an acid action. When first sublimed it occurs in transparent colourless vitreous masses. These masses are remarkable from becoming opaque upon exposure to the air. The fact is, that arsenious acid is dimorphous; *i. e.*, capable of assuming two different crystalline forms, each belonging to different systems (see Chap. X.), and this change in the opacity is probably owing to this property. As sold in the shops arsenious acid is a fine, white powder. At 380° Fah. it volatilises without previously melting. In this process it produces a white smoke, and the acid condenses unchanged upon cold surfaces; and if the sublimation be slowly carried on in a glass tube, transparent octohedral crystals of it are obtained.

To the taste arsenious acid is slightly sweetish; and the idea formerly prevalent that it is acid is a mistake. It feebly reddens litmus paper, and combines with bases, forming *arsenites.*

This compound is so frequently taken and administered as a poison, that the tests for it become of very great consequence. They will be found after the account of the sulphuret.

3. Arsenic Acid.

This is prepared by heating arsenious and nitric acid, evaporating, and then again heating in a crucible the compound thereby produced. It has a sour, metallic taste, reddens vegetable blues, and forms with bases salts called *arsenates.* It is a powerful poison.

4. Sesquichloride of Arsenic.

When arsenic is thrown into a jar of chlorine it takes fire, and the sesquichloride is formed. It is a colourless, volatile liquor, which fumes very much when exposed to the air, and is hence called the fuming liquor of arsenic. Mixed in any quantity with water it is resolved into arsenious and hydrochloric acids.

5. Protosulphuret of Arsenic.

This compound occurs in the mineral kingdom, and is known in such by the name of realgar. It may be artificially prepared by heating together arsenic and the sesquisulphuret. It is a ruby-red, crystalline mass.

6. SESQUISULPHURET OF ARSENIC.

This also is a mineral, and is called orpiment. It is prepared by transmitting a stream of sulphuretted hydrogen through a solution of arsenious acid. Orpiment has a rich yellow colour, and is the foundation of the pigment called king's yellow. It is soluble in alkalis.

An impure orpiment is sold in the shops for the purpose of killing flies.

TESTING FOR ARSENIC.

Owing to the frequency with which arsenious acid is used for the purpose of poisoning, it is very important to be able to decidedly detect it. Hence the tests for it are of great interest and importance. These may be conveniently arranged under—1. the liquid tests, including in this division sulphuretted hydrogen; 2. the reduction test; 3. Marsh's test; and, 4. Reinsch's.

1. FLUID TESTS.

a. If to a solution containing arsenious acid a solution of nitrate of oxide of silver with a little ammonia be added, a canary-yellow arsenite of oxide of silver is thrown down. The reason that a little ammonia must be added, is that the nitric acid and oxide of silver will not separate when brought into contact with arsenious acid, unless there be an alkali to combine with the nitric acid. The nitrate of ammonia is retained in solution.

It happens, however, that arsenite of oxide of silver is very soluble in ammonia; so that if too much ammonia be added, that which does not combine with the nitric acid dissolves the arsenite of oxide of silver, and there is no yellow precipitate. To obviate this source of fallacy, a compound salt (the ammoniacal nitrate of silver) is used which contains the exact quantity of ammonia that is required and no more.

b. If to a solution of arsenious acid a solution of sulphate of oxide of copper, to which a little ammonia is added, be dropped, the bright green arsenite of oxide of copper, or Scheele's green, is precipitated.

c. If a current of sulphuretted hydrogen be transmitted through a solution of arsenious acid the yellow sesquisulphuret of arsenic is precipitated. In case any pure alkali be present, which would dissolve the orpiment, a little hydrochloric acid should be added. This of course forms a neutral salt with the alkali.

2. The Reduction Test.

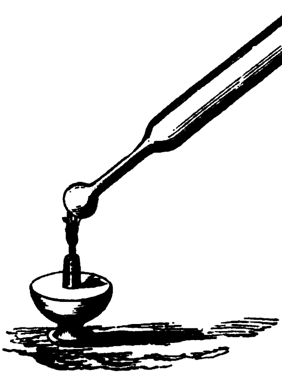
If a small quantity of arsenious acid be introduced into the bottom of a test tube, as *fig.* 42., and gently heated in the flame of a small lamp, the acid is vaporised, ascends in the tube, and condenses in brilliant crystals, which, if perfect, are octohedrons. (See Chap. X.) If perfect crystals are not obtained, they

Fig. 42.

may be observed to have equilateral triangular faces. No other white powder volatilises and condenses in such crystals.

Another small quantity of arsenious acid should be mixed with black flux (*i. e.*, a mixture of charcoal and carbonate of potassa), and placed in the bottom of a tube shaped like *fig.* 43. In order to prevent any of it adhering to the sides it should be introduced on a small slip of paper. The tube must then be very gently heated by passing it backwards and forwards over a spirit lamp, and the moisture that is thereby condensed in the inside carefully removed by bibulous paper. The bulb is then held in the flame until it has attained a red heat; upon which the oxygen of the arsenious acid and the carbon of the flux combine together, and the metallic arsenic rises in vapour and condenses in the narrow neck above the bulb. Here it forms a metallic crust of an iron-grey colour externally, and crystallises on the inner

Fig. 43.



surface. By cutting off the narrow neck with a file, and heating it in the tube *fig.* 42., arsenious acid can be reprocured.

The action of the alkali in the above is to retain the arsenious acid until the temperature is sufficiently high for its decomposition.

It was stated above that sulphuretted hydrogen threw down with arsenious acid yellow sulphuret of arsenic. It was purposely omitted to be stated then that the sulphurets of antimony, tin, cadmium, and selenium are also yellow. In order to decide if the precipitate contain arsenic, it is necessary to heat some of it* in a tube like that in *fig.* 43. along with black flux, and if arsenic be present it is sublimed and deposited in the metallic form. In this instance it is the alkali that effects the decomposition, its potassium uniting with the sulphur to liberate the arsenic.

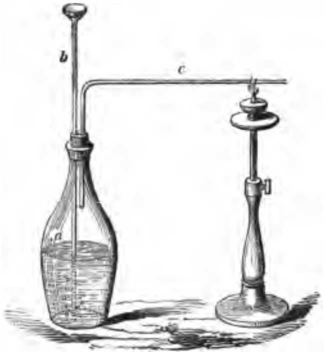
* Previously well dried.

If arsenious acid be obtained as a white powder, or in a colourless solution, the above tests will readily detect its presence; but if it be dissolved in coloured solutions, as tea or porter, the liquid tests are of no value, as the colour of the precipitates is interfered with. The two following plans provide for such emergencies.

3. Marsh's Test.

This consists in combining hydrogen with the arsenic, burning the arseniuretted hydrogen thereby produced, and thus obtaining a crust of metallic arsenic. It may be done in the apparatus represented in *fig.* 44. The jar *a* contains the mixed fluid, holding in solution arsenious acid, to which *pure* zinc is added. Sulphuric acid is poured down the funnel *b*, and the gas which is evolved through *c* is heated by a spirit lamp. The metal separates, and forms in a crust a little beyond the place where the flame is applied.

Fig. 44.

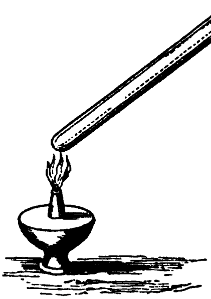


Anti-moniuretted hydrogen burns with a similar flame to that of arseniuretted hydrogen, and metallic antimony is deposited. The two metals, however, may be readily distinguished; for while, as has just been stated, the arsenic, when heated, becomes converted into arsenious acid, heat merely converts the antimony into a white powder, which is the sesquioxide of antimony.

4. Reinsch's Test.

This consists in boiling the solution of arsenic along with hydrochloric acid and copper wire for some time. The copper and the arsenic then combine, and the wire acquires a dark colour. If these are dried, and introduced into a narrow test tube, as *fig.* 45., and heated, the arsenic separates, rises in vapour, and forms a ring. This arsenic may be converted into arsenious acid in the manner before

Fig. 45.



mentioned; and this arsenious acid may be further tested by the ammoniated nitrate of silver, ammoniated sulphate of copper, and sulphuretted hydrogen.

Arsenic acid forms a reddish-brown precipitate of arseniate of silver with a solution of nitrate of silver. Metallic arsenic, or arsenious acid, may be converted into arsenic acid by the action of nitric acid, and then this test applied.

In testing organic mixtures supposed to contain arsenic, various details are necessary to be attended to. These will be found described in books upon legal medicine.

XXIII. TELLURIUM AND URANIUM.

These two very rare metals are of little importance.

(E.) XXIV. COPPER.

This frequent metal was well known to the ancients. It derives its name from having been originally obtained from the island of Cyprus. It is the Venus of the alchemists. It is said to have been detected in a few plants; but not in those that form the food, at least to any extent, of animals. Its most abundant ores are the sulphurets,—the copper pyrites and grey copper ore of the miners. Both of these contain also iron.

Owing to the copper pyrites being a double sulphuret, the extraction of the copper is a tedious and somewhat difficult process. First of all the ore is roasted, and a portion of the sulphur in that way expelled. The sulphur that remains is oxidated, and becomes sulphuric acid, and then the mass that remains in the furnace consists of sulphates, of oxides of iron and copper. This mixture is heated along with fresh ore and silicious matter, and a great part of the oxide of iron unites with the silicic acid and rises to the surface in the form of slag, and the copper becomes again a sulphuret. This is roasted, converted into an oxide, and then reduced by heating it with carbonaceous matter.

Another valuable ore of copper is the carbonate. It owes this value to the ease with which it is reduced.

Copper is the only metal that is of a red colour. It is ductile and malleable, and only inferior in tenacity to iron. It is hard, elastic, and sonorous. It is difficult of fusion. Its specific gravity is nearly 9. Cast in a mould it does not take a sharp impression, and hence copper coins require to be stamped.

Exposed to air and moisture its surface rusts, and the oxide thus formed unites with the carbonic acid of the atmosphere, and the green carbonate of the protoxide is formed. At a red heat the black protoxide is formed.

The atomic weight of copper is 31, and its symbol Cu.

The compounds of copper described in this section are as follow:—

								Formula.
1.	*Dinoxide, or suboxide of copper*	Copper	2 atoms	62+Oxygen	1 atom	8=70		$2Cu+O$.
2.	*Protoxide of copper*	Copper	1 atom	31+Oxygen	1 atom	8=39		$Cu+O$.
3.	*Dichloride of copper*	Copper	2 atoms	62+Chlorine	1 atom	35=97		$2Cu+Cl$.
4.	*Chloride of copper*	Copper	1 atom	31+Chlorine	1 atom	35=66		$Cu+Cl$.
5.	*Disulphuret of copper*	Copper	2 atoms	62+Sulphur	1 atom	16=78		$2Cu+S$.
6.	*Sulphuret of copper*	Copper	1 atom	31+Sulphur	1 atom	16=47		$Cu+S$.

1. DINOXIDE, OR SUBOXIDE OF COPPER.

This occurs native, forming beautiful ruby-red, transparent, octohedral crystals. It may be formed artificially in a variety of manners. For instance, if five parts of the protoxide are heated with four of copper filings it is procured; or if a solution of sulphate of protoxide of copper be boiled along with honey, sulphate of potassa is formed, and the honey appropriates the half of the oxygen of the protoxide, upon which the dinoxide falls down as a brown-red powder.

This dinoxide has not a very strong affinity for acids, but still it does form salts, most of which are colourless, but which are very apt to absorb oxygen and to pass into compounds of a higher base. It tinges glass, when heated with that substance, a ruby-red; but this colour is apt to pass into a green, owing to its becoming converted into protoxide, which stains glass that colour.

2. PROTOXIDE OF COPPER.

This is often called the black oxide. It occurs native; in which case it constitutes the copper-black of mineralogists. It may be artificially prepared by heating sheet-copper to a red heat in contact with the air, or by calcining nitrate of potassa, or by precipitating it from its salts by means of pure potassa.

It is of a dark colour, not acted upon by heat alone; but if mixed with carbonaceous matter readily reduced to a metallic state. It forms salts with almost all acids, and all these salts have a green or blue colour. It is also soluble in ammonia, and forms with it a blue colour; a property which distinguishes it from all other substances.*

Salts of the protoxide, besides striking this characteristic blue with ammonia, give a dark-brown precipitate with sulphuretted hydrogen. On the addition of an alkali to this the hydrated oxide, which becomes (it must be boiled, however, with water) black, copper in the metallic state may be obtained from a solution of them by introducing a rod of iron or zinc. Ferrocyanide of potassium gives a peculiar mahogany-coloured precipitate with them.

3. Dichloride of Copper.

If copper filings be introduced into chlorine the metal takes fire, and both the dichloride and chloride are formed. The dichloride may be prepared separately by heating copper filings with twice their weight of bichloride of mercury. It was prepared in this manner by Boyle, and was termed by him, from its resemblance to resin, resin of copper. If, also, the solution of chloride of copper is kept in a bottle in contact with metallic copper, the dichloride is deposited in crystalline grains. Its colour is white, yellow, or brown, according to the mode in which it is prepared. It is very apt to absorb oxygen from the air, and form a green compound of oxide and chloride of copper.

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* The oxide alone is not soluble in ammonia, and it must be combined with an acid. *See* Ammoniacal Salts in Book III.

4. CHLORIDE OF COPPER.

This is prepared by dissolving protoxide of copper in hydrochloric acid. Water is formed, and green crystals of hydrated chloride of copper are thrown down. When heated, they fuse, give off the water, and become brown. If the heat, however, exceed 400°, half the chlorine is lost, and the dichloride formed.

Chloride of copper is soluble in alcohol. If to this solution an alkali be added, an oxychloride $(Cu+Cl)+3(Cu+O)+(OH)$ is formed. This is the pigment named Brunswick green.

5. DISULPHURET OF COPPER.

This may be artificially formed by heating copper filings with a third of their weight of sulphur. The combination is attended with such evolution of heat that the mass becomes luminous.

6. SULPHURET OF COPPER.

This is precipitated from a solution of any salt of the protoxide by means of sulphuretted hydrogen. It has a dark-brown colour.

XXV. Lead, or Plumbum.

This metal has been known from the remotest antiquity. It is the Saturn of the alchemists.

A great many ores of lead occur in nature; but for commercial purposes lead is obtained from the sulphuret, or *galena* of mineralogists. This is first oxidated, and then reduced to the metallic state by the action of carbonaceous matter.

Lead is a bluish-grey metal, the surface of which, when recently cut, has a brilliant metallic lustre. Its surface, however, soon tarnishes by exposure to the air. It is heavy, having a specific gravity of more than 11. It is malleable and ductile, but possesses little tenacity. It fuses at 612°.

At high temperatures lead oxidates very rapidly.

In recently boiled distilled water, and preserved in close vessels, lead undergoes no change; but if air be present it is rapidly oxidated, and this oxide of lead combines with the carbonic acid present in the air, and carbonate of oxide of lead is produced. As this salt is poisonous, and, to a certain extent, soluble in water, that fluid is apt to acquire poisonous properties when kept in leaden cisterns or conveyed along leaden pipes. If, however, the water, as almost all spring water does, contain sulphates and chlorides, then a sulphate or chloride of lead is formed; and, as these are quite insoluble, they incrust the lead, and not only do no harm to the water, but protect the lead from being farther acted upon.

The atomic weight of lead is 103, and its symbol Pb.

The compounds of lead described in this section are as follow: —

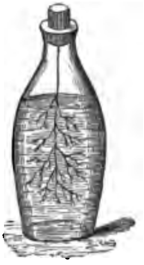
				Formula.
1.	*Protoxide of lead*	-	Lead 1 atom 103+Oxygen - 1 atom 8=111	$Pb+O$.
2.	*Peroxide of lead*	-	Lead 1 atom 103+Oxygen - 2 atoms 16=119	$Pb+2O$.
3.	*Red oxide of lead*	-	A mixture of the two preceding oxides.	
4.	*Chloride of lead*	-	Lead 1 atom 103+Chlorine 1 atom 35=138	$Pb+Cl$.
5.	*Iodide of lead*	-	Lead 1 atom 103+Iodine - 1 atom 126=229	$Pb+I$.
6.	*Sulphuret of lead*	-	Lead 1 atom 103+Sulphur 1 atom 16=119	$Pb+S$.

1. PROTOXIDE OF LEAD.

This may be prepared by exposing lead, while melted, to a current of hot air. If it is not allowed to fuse it is of a pale yellow colour, and is the massicot of old writers. When fused or semi-vitrified, it is of a red colour, and is called litharge. It is used largely in the arts, principally to mix with glass, which it renders more brilliant and fusible, and for glazing earthenware. It is the base of all hair dyes.

The protoxide is the only oxide of lead that forms salts. Any of these salts may be recognised by giving a black precipitate (of sulphuret) with sulphuretted hydrogen; by being precipitated as a white hydrate by pure alkalis, and which hydrate is soluble in an excess of the alkali; as a white carbonate by alkaline carbonates; as a white sulphate by soluble sulphates; as a yellow iodide by iodide of potassium; and as yellow chromate (chrome yellow) by bichromate of potassa.

Fig. 46.



Lead also is separated from its salts in a metallic state by iron or zinc. If an ounce of acetate of lead be dissolved in a wine-quart of water, and a piece of zinc suspended in the solution by means of a thread, the lead is deposited in an arborescent form, and gives rise to what has been called *arbor Saturni*, *fig.* 46.

2. Peroxide of Lead.

This is obtained by adding dilute nitric acid to red lead. The acid unites with the protoxide and liberates the peroxide. It is a puce-coloured powder. It does not unite with acids.

3. Red Oxide of Lead.

This is the minium of commerce. It is prepared by exposing massicot at a temperature of about 600° to a current of air. It then slowly absorbs oxygen, and a portion of it becomes converted into the peroxide. It is of a red colour, is used as a pigment, and also by the glass manufacturers.

4. Chloride of Lead.

This occurs native, and may be formed by adding chloride of sodium to any soluble salt of lead. It is a white crystalline powder, which, when heated, fuses, and then forms a gelatinous-like mass sometimes called horn lead.

5. Iodide of Lead.

This is formed by adding iodide of potassium to acetate of lead, the result being acetate of potassa and yellow iodide of lead.

6. SULPHURET OF LEAD.

This may be artificially formed either by the action of sulphuretted hydrogen on a solution of a lead salt, or by fusing together sulphur and lead. It is a black, insoluble powder.

* * *

XXVI. BISMUTH.

This somewhat rare metal was known to the ancients, although it was confounded by them with lead or tin. Its name bismuth or wismuth was conferred upon it by the German miners. Unimportant ores of it occur, but its principal commercial source is the native bismuth.

It is a brittle metal of a white colour with a reddish tinge. Its density is about 10. When hot it may be hammered into plates. At 470° it fuses and at a red heat it sublimes.

Bismuth is easily oxidised by the action of nitric acid and by fusion in open vessels; and at its subliming point it burns with a bluish flame, emitting copious fumes of its oxide. It is used in the arts.

The atomic weight of bismuth is 71, and its symbol Bi.

The only compound of bismuth described in this section is as follows:—

					Formula.
Protoxide of Bismuth -	Bismuth -	1 atom 71	+ Oxygen -	1 atom 8 = 79	$Bi+O$.

This may be prepared either by heating bismuth in the open air or by the action of potassa upon the nitrate of bismuth. It has a yellow colour. It forms salts with acids. These are known by the brown precipitate which they give with sulphuretted hydrogen (the sulphuret of bismuth), and the brown iodide by iodide of potassium. A strong solution of a salt of bismuth is decomposed by being added to water; the water retains a great part of the acid in solution, and a salt with only one atom of acid to three of base is precipitated.

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(F.) XXVII. SILVER, OR ARGENTUM.

Silver has been known from the most remote periods of antiquity. It is the Diana or Luna of the alchemists.

It occurs in nature either native or as a constituent of various metallic ores. Its extraction from these latter is managed in two manners, both of which will be noticed in Book V.

The silver of commerce, however, commonly contains both copper and gold. For chemical purposes it may be obtained by dissolving the commercial metal in nitric acid, and precipitating the solution thus obtained by chloride of calcium. The chloride of silver thus thrown down is then reduced by white flux, that is, carbonate of potassa; andt he chlorine combines with the potassium, and the silver is thus obtained pure.
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Silver is a very brilliant metal, and of all metals is the whitest. It takes a very high polish, and is very malleable and ductile, and possesses considerable tenacity. When pure it is so soft that it may be cut with a knife, and what is commonly called silver in coins, or articles for domestic use, is in fact a mixture, or alloy of silver with copper.

The specific gravity of silver is more than 10. It fuses at about 1870°. Neither air nor water have, properly speaking, any action upon it; but when it is kept melted for a considerable period it absorbs oxygen to the amount indeed of more than twenty times its volume; but when it becomes solid it parts with the whole of this, and, by discharging thus the oxygen, its surface is thrown into a peculiar appearance known as that of granulated silver. Also, silver, when exposed to the air, tarnishes; but this is owing to the action of the sulphuretted hydrogen contained in the atmosphere, a sulphuret of silver being formed.

The atomic weight of silver is 108, and its symbol Ag.

The compounds of silver noticed in this section are as follow:—

				Formula.
1. *Oxide of silver*	- Silver 1 atom 108+Oxygen	- 1 atom	8=116	Ag+O.
2. *Chloride of silver*	- Silver 1 atom 108+Chlorine	1 atom	35=143	Ag+Cl.
3. *Iodide of silver*	- Silver 1 atom 108+Iodine	- 1 atom	126=234	Ag+I.
4. *Sulphuret of silver*	- Silver 1 atom 108+Sulphur	- 1 atom	16=124	Ag+S.

1. OXIDE OF SILVER.

This may be obtained in the state of a hydrate by adding lime-water to a solution of nitrate of silver.

It is a brown powder, very prone to resolve itself into oxygen and metallic silver. It forms salts with acids. It also forms a curious compound with ammonia, that will be hereafter noticed.

Salts of the oxide of silver are known by giving a black precipitate with sulphuretted hydrogen, and a white one with hydrochloric acid; which latter dissolves in ammonia, but not in nitric acid.

A number of metals, as zinc or iron, precipitate metallic silver from solutions of its salts. In this manner a tree like the lead tree may be formed, which has been called *arbor Dianæ*, or *arbor lunæ*.

2. Chloride of Silver.

This occurs native, in which case it is called horn silver, and may readily be prepared by a solution of common salt with nitrate of silver. It is a curdy precipitate, insoluble in acids, but quite so in ammonia. Like all the salts of silver it is remarkably sensitive to light, and when in contact with organic matter becomes black under their influence. Hence its use in photography.

3. Iodide of Silver.

This is obtained by decomposing any soluble salt of silver by iodide of potassium. It is a primrose-yellow substance, insoluble in water, and nearly so in ammonia. It fuses readily, and is even more sensitive to the light than the chloride. In consequence of this property it is the basis of the impression in the Daguerréotype process.

4. SULPHURET OF SILVER.

This occurs native, constituting the silver glance of mineralogists, and may be artificially formed by heating sulphur and silver together. It is remarkable for being soft and malleable.

* * *

XXVIII. MERCURY, OR HYDRARGYRUM.

This metal is still commonly known by its alchemical name. Another very common appellation of it is quicksilver. It is one of the metals that the ancients were acquainted with. It is a metal of comparatively rare occurrence; it is found in small quantities, however, native, but the ore from which the mercury of commerce is extracted is the bisulphuret, or cinnabar.

The mode of extracting mercury from cinnabar is by heating it along with lime, the result of the process being the metal and sulphuret of calcium. Another method consists in roasting the ore, by which the sulphur is converted into sulphurous acid and the mercury is volatilised. Commercial mercury usually contains a small admixture of other metals, from which it may be separated by distillation.

Mercury at ordinary temperatures is fluid. It is without odour or taste, and has a whitish colour that somewhat resembles that of tin. It becomes solid at about 40° below zero, and is then malleable and can be cut with a knife. At 662° it vaporises and condenses again on cool surfaces in metallic globules. Its specific gravity is more than 13.

Mercury if quite pure does not tarnish on exposure to the air; but if it contain other metals, these rapidly oxidate, in consequence of a galvanic action being set up between the metals. Nitric acid acts upon it, oxidates it, peroxide of nitrogen being evolved, and nitrate of oxide of mercury being formed. Sulphuric acid with the aid of heat forms with it sulphate of oxide of mercury.

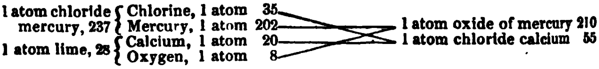
The atomic weight of mercury is 202 *, and its symbol Hg.

The compounds of mercury described in this section are as follow:—

			Formula.
1. *Protoxide of mercury* -	Mercury 1 atom 202+Oxygen - 1 atom	8=210	$Hg+O$.
2. *Peroxide of mercury* -	Mercury 1 atom 202+Oxygen - 2 atoms	16=218	$Hg+2O$.
3. *Protochloride of mercury*	Mercury 1 atom 202+Chlorine 1 atom	35=237	$Hg+Cl$.
4. *Per or bichloride of merc.*	Mercury 1 atom 202+Chlorine 2 atoms	70=272	$Hg+2Cl$.
5. *Iodide of mercury* -	Mercury 1 atom 202+Iodine - 1 atom	126=328	$Hg+I$.
6. *Biniodide of mercury* -	Mercury 1 atom 202+Iodine - 2 atoms	252=454	$Hg+2I$.
7. *Sulphuret of mercury* -	Mercury 1 atom 202+Sulphur - 1 atom	16=218	$Hg+S$.
8. *Bisulphuret of mercury*	Mercury 1 atom 202+Sulphur - 2 atoms	32=234	$Hg+2S$.

1. Protoxide of Mercury.

This is often called the black oxide. It is prepared by the action of lime upon chloride of mercury. Double decomposition takes place.



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* Many state it 101.

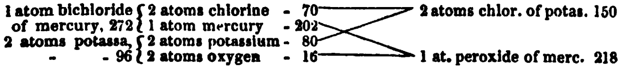
If mercury be rubbed for a long time with fatty matter, the metal disappears, and is technically said to be killed. A portion of it is perhaps oxidated.

Protoxide of mercury is black in colour, without taste or smell, is insoluble in water and the alkalies, but dissolves in nitric acid. It is readily decomposed by light, and is resolved into mercury and the peroxide. It forms salts; the best tests for which are sulphuretted hydrogen, which throws down the black protosulphuret, and hydrochloric acid, which precipitates the insoluble white protochloride, or calomel.

2. PEROXIDE OF MERCURY.

This is frequently called red precipitate. It may be prepared in a variety of manners. One consists in vaporising mercury in an apparatus for the purpose, and exposing it for a very long period to the air. This process occupies, or did occupy, many weeks, and the apparatus in which it was effected was called "Boyle's hell," from the idea that the metal was tortured in it.

A much more convenient plan is to decompose the bichloride of mercury by means of potassa. The decomposition that takes place is shown in the following table:—

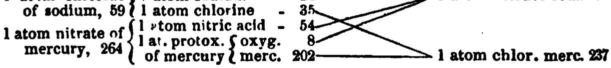


Prepared in this latter manner, this oxide is of an orange-red colour, which becomes black with heat. It is slightly soluble in water, forming a solution that has an acrid taste. Heated to redness it separates into oxygen and mercury. It combines with acids and forms salts. These may be distinguished from salts of the black oxide by giving with ammonia a white precipitate, which is a double salt.

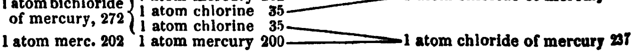
Any solid compound of mercury, if heated with carbonate of potassa, gives a sublimate of metallic mercury.

3. Protochloride of Mercury, or Calomel.

This mercurial compound has received a vast variety of names: manna of the metals, sweet mercury, calomel, submuriate of mercury, and many others. It occurs (but very rarely) native, and constitutes the horn-silver of mineralogists. It may be artificially prepared either by precipitation or sublimation. It is conveniently obtained in the former manner by the action of chloride of sodium upon the nitrate of the protoxide of mercury. The change that takes place is as follows:—



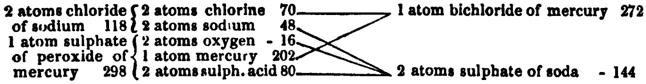
To obtain calomel by sublimation, four parts of the bichloride of mercury are rubbed up with three parts of metallic mercury, and the mass sublimed. The union that takes place is obvious.


1 atom bichloride of mercury, 272
1 atom mercury 202
1 atom chlorine 35
1 atom chlorine 35
1 atom merc. 202
1 atom mercury 200
1 atom chloride of mercury 237
1 atom chloride of mercury 237

As prepared by sublimation, calomel is in semitransparent cakes; but when it is procured by precipitation it is a white powder. It is volatilised by heat, and during this process a portion is always converted into bichloride and mercury. It is not poisonous, has no taste, and it is exceedingly insoluble in water. These three negative properties at once distinguish it from the bichloride or corrosive sublimate.

4. BICHLORIDE OF MERCURY, OR CORROSIVE SUBLIMATE.

If mercury be heated in chlorine it takes fire, and the very virulent poison bichloride of mercury is formed. For medicinal purposes it is prepared by mixing together one portion of sulphate of peroxide of mercury with two of chloride of sodium. The following decomposition takes place:—


2 atoms chloride of sodium 118
1 atom sulphate of peroxide of mercury 298
2 atoms chlorine 70
2 atoms sodium 48
2 atoms oxygen - 16
1 atom mercury 202
2 atoms sulph. acid 80
1 atom bichloride of mercury 272
2 atoms sulphate of soda - 144

It is a semi-transparent, colourless substance of a crystalline texture. It has an acrid, burning, and metallic taste; is soluble in water, and is extremely poisonous. It forms with the organic principle albumen an insoluble compound, and this substance is therefore an antidote to it.

Corrosive sublimate may be determined to be a permercurial salt, *a*, by adding to a solution of it caustic ammonia, when a white powder is thrown down: *b*, by adding protochloride of tin, upon which calomel is precipitated; and if the salt of tin be added in excess, the calomel becomes grey, owing to the separation of mercury in the metallic form: *c*, by sulphuretted hydrogen in excess throwing down the black bisulphuret; but if the sublimate be in excess, the white chlorosulphuret: and, *d*, by the galvanic test, *fig.* 47. A drop of the solution is placed upon a sovereign and a key is made to touch the gold and the solution; upon which the metallic mercury attaches itself to the gold.

Fig. 47.

5. Iodide of Mercury.

This may be formed by triturating iodine and mercury together. It is an olive-green powder.

6. Biniodide of Mercury.

This may be prepared by expelling, by means of heat, an atom of mercury from the previous compound. It is a scarlet-red powder. Heated, it fuses, and sublimes in crystals that are first yellow, but become scarlet-red. Sometimes, if kept very still, they retain their yellow colour; but the slightest touch causes them to become red.

7. Sulphuret of Mercury.

This is an unimportant compound, of a black colour.

8. BISULPHURET OF MERCURY.

When a solution of sulphuretted hydrogen in excess is added to a solution of corrosive sublimate, this compound is precipitated, and is then black. If sublimed, however, it becomes red. Or it may be prepared by mixing sulphur and mercury together, and subliming. It is then often called factitious cinnabar, and, when powdered, the beautiful pigment vermilion. Bisulphuret of mercury occurs native, and is called cinnabar.

* * *

XXIX. GOLD, OR AURUM.

Gold has been known from the remotest antiquity, and in all ages seems to have commanded the homage and respect which it receives in our day. By the alchemists it was termed Jupiter, or Rex metallorum.

There are no ores of gold, but it occurs in the metallic state either alone or combined with other metals. It is found either in primitive or metamorphic rocks, or in small particles or gold dust in the beds of the streams that flow from them. Until recently its principal source was South America; but, as the whole world knows, it has lately been abundantly found in California and Australia.

The separation of gold from other metals, and from sand, will be shortly considered in Book V.

Gold is the only metal that has a yellow colour. Of all metals it is the most malleable and ductile. Its density is 19.

Gold may be exposed to air and moisture for any period without being oxidated. Neither does it combine with oxygen when fused in the open air. No ordinary acid has any effect upon it; but if immersed in a mixture of nitric and hydrochloric acids it combines with a portion of the chlorine of the latter.

The atomic weight of gold is 199, and its symbol Au.

The compounds of gold noticed in this section are as follow:—

						Formula.
1.	*Protoxide of gold* -	Gold 1 atom	199+Oxygen	1 atom	8=207	$Au+O$.
2.	*Peroxide of gold* -	Gold 1 atom	199+Oxygen	3 atoms	24=223	$Au+3O$.
3.	*Protochloride of gold*	Gold 1 atom	199+Chlorine	1 atom	35=234	$Au+Cl$.
4.	*Perchloride of gold* -	Gold 1 atom	199+Chlorine	3 atoms	105=304	$Au+3Cl$.

1. Protoxide of Gold.

This is obtained by the action of potassa on a solution of protochloride of gold. It has a green colour, and soon after its preparation it spontaneously changes into metallic gold and peroxide.

2. Peroxide of Gold.

This is prepared by the action of alkalies upon the perchloride, the result of this mixture being an alkaline chloride and this oxide. When freed from water this has a black colour. It does not combine with acids, but apparently forms salts with alkalies. Hence it has been proposed to call it auric acid.

3. PROTOCHLORIDE OF GOLD.

This may be procured by depriving the terchloride of part of its chlorine by heat.

4. PERCHLORIDE OF GOLD.

This is the only important chemical compound of gold. It is formed by dissolving gold in nitro-hydrochloric acid, and then evaporating. As obtained in this manner, it is in crystalline needles of an orange-red colour, and a styptic, disagreeable taste. It is soluble in water, ether, and alcohol. The ethereal solution is used for gilding, as it deposits metallic gold on a number of substances to which it is applied, particularly under the influence of light.

Gold is precipitated from its solutions by a number of substances. The one usually employed as a test for it is sulphate of the protoxide of iron.* Protochloride of tin throws down, when added to a golden solution, the purple of Cassius, noticed under tin.

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XXX. PLATINUM.

This is a valuable, but rare metal. It only occurs in the metallic state associated with other metals. Its principal sources are South America and the Uralian mountains.

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* The results are sesquisulphate of the peroxide of iron, perchloride of iron, and the *metallic gold.*
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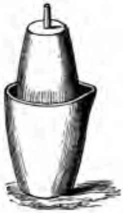
Fig. 48.



Platinum has a white colour, very like that of silver, but of inferior lustre. Of all metals it is the heaviest; its specific gravity being 21·5. It is a soft metal, and, like iron, two pieces of it, at a high temperature, can be *welded* together. Neither air, nor moisture, nor the highest heat of the forge of the smith, have any effect upon it. None of the pure acids have any action on it, and only nitro-hydrochloric acid dissolves it. Caustic potash, however, if melted upon it, oxidates it. These two properties of resisting reagents and bearing so high a temperature, render it most useful to the chemist, who constantly employs crucibles (*fig.* 48.) made of it in the performance of analyses.

Platinum heated before the oxy-hydrogen blowpipe fuses and burns.

There are two very curious modifications of platinum, called spongy platinum and platinum black. The former is prepared by mixing a solution of chloride of platinum in alcohol with an alcoholic solution of hydrochloride of ammonia. A double chloride is precipitated as a fine yellow powder. This is immediately rolled up into little balls and exposed to the fire. All the constituents of the compound salt, save the platinum, are thereby expelled, but the metal that is left has a light, spongy appearance. This spongy platinum possesses the remarkable property of causing oxygen and hydrogen to unite. This property was mentioned at p. 103., and there also a diagram of Dobereiner's lamp, founded upon this fact, is given.

Platinum black is the metal in a state of still finer division. Its power of promoting the union of hydrogen and oxygen is still greater, and advantage is taken of its properties in this respect in many processes of organic chemisty.

The atomic weight of platinum is 99, and its symbol Pl.

The compounds of platinum described in this section are as follow:—

				Formula.
1. *Protoxide of platinum*	Platinum 1 atom 99+Oxygen	1 atom	8=107	$Pl+O$.
2. *Peroxide of platinum*	Platinum 1 atom 99+Oxygen	2 atoms	16=115	$Pl+2O$.
3. *Protochloride of platinum*	Platinum 1 atom 99+Chlorine	1 atom	35=134	$Pl+Cl$.
4. *Perchloride of platinum*	Platinum 1 atom 99+Chlorine	2 atoms	70=169	$Pl+2Cl$.

1. Protoxide of Platinum.

It is produced by digesting the protochloride with as much potassa as is exactly sufficient for the decomposition. If an excess of potassa be added it dissolves it. This oxide forms salts with acids.

2. Peroxide of Platinum.

This is prepared by decomposing the sulphate of the peroxide of platinum by means of soda. It is with difficulty obtained in a pure form, as, like the peroxide of gold, it has a tendency to act as an acid and combine with the alkali.

3. Protochloride of Platinum.

This is formed by expelling an atom of chlorine from the bichloride by heat. It is a greenish-olive powder.

4. Perchloride of Platinum.

This is obtained by digesting metallic platinum in nitro-hydrochloric acid and then evaporating. It is soluble in water and alcohol. It is used as a test for potassa.

A soluble salt of platinum as this bichloride may be detected by giving, with hydrochlorate of ammonia, a yellow precipitate, which, when heated, becomes spongy platinum; a substance that is insoluble in all acids save nitro-hydrochloric acid.

* * *

XXXI. Palladium, Iridium, Rhodium.

These rare metals are associated with platinum, and are of no practical importance.

CHAP. VIII.

ALLOYS AND AMALGAMS.

THE metals would appear to unite with one another in almost all proportions, in the same manner as sulphuric acid and water do. Hence, metallic combinations with one another, or alloys and amalgams, as they are named, are not chemical compounds, but mechanical admixtures: nevertheless, the properties of alloys often differ very considerably from those of the metals that form them; and the economical uses, and the prospect of still greater economical uses, to which they may be put, render them very interesting compounds.

In nature a great many alloys exist; indeed, it is rare to fine a metal alone; but very curiously, with the exception of the alloy found in meteoric stones, which the Esquimaux employ in the construction of their knives, no native alloy is of any practical use. In practice, again, we find that, while some metals may be used advantageously in a pure state, as platinum, iron, lead, copper, and zinc, others must (although in common language we regard them as pure) be alloyed. Thus, in the arts, gold, silver, and tin are always mixed with other metals; and if this were not done, would be so soft as to be comparatively useless; antimony and bismuth, again, must be alloyed with other metals to prevent their excessive brittleness; and some of the more frequent alloys possess properties that no simple metal has, and save iron, one, perhaps, of the greatest importance in the arts.

In fact, as far as regards its properties and uses, every separate alloy must be taken as a distinct metal. More than sixty alloys have been, more or less, studied, but a great many more may be formed; and there can be little doubt but that very desirable results would attend a searching examination of such. In this chapter we can only notice the composition and properties of a very few.

I. Brass.

This is an alloy of copper and zinc, and has been known from the remotest antiquity. It was formerly prepared by mixing together metallic copper, calamine (*i. e.*, carbonate of zinc), and charcoal, and exposing the mixture to bright ignition; but for many years it has been prepared by the fusion together of the two metals. The best proportion of the two elements for making brass for ordinary purposes is said to be two parts of copper to one of zinc.

Brass, when cold, is malleable and ductile, and may be beat into leaves and drawn into wire. When heated, however, it becomes brittle. It has a bright yellow colour, takes a fine polish, is little acted upon by the air, and is easily worked with turning and other instruments. In the state of wire it is much used in the manufacture of pins, and, beaten into leaves, it constitutes the imitation of gold-leaf called Dutch gold.

Prince's metal is a variety of brass in which the zinc preponderates in quantity over the copper. The *platin*, or button-metal of Birmingham, is another mixture of brass and zinc; and pinchbeck, an alloy formerly a good deal used for making cheap watch-cases, is brass containing a large excess of copper over the zinc.

II. BRONZE AND GUN METAL.

Bronze essentially consists of an alloy of copper and tin, although a little lead and zinc are often added. Before the discovery of the art of working iron, bronze seems to have been used for the manufacture of weapons and of ordinary implements. Owing to its fusibility and hardness, bronze is now much used for making cannon, and also, as indeed has been the custom from time immemorial, for the casting of statues.

Bronze and gun-metal contain about eight parts of copper to one of tin. Somewhat curiously, this was the proportion fixed upon by the pre-historic manufacturer of bronze swords, &c.

III. BELL METAL.

Common bell metal contains about 100 parts of copper to from 40 to 60 of tin. This alloy is brittle, but very sonorous; and hence the use to which it is put. The metal of house-clock bells contains more tin than that cast for church bells, and the metal of a repeater has, in addition, a little zinc.

A very fine kind of bell-metal is made by the Chinese for the purpose of forming gongs, and called tamtam. Tamtam contains less tin than our bell metal.

IV. Speculum Metal.

This is a white, brittle alloy, also composed of copper and tin; and which, owing to the fine polish that it will take, is used for making the mirrors of reflecting telescopes.

It is made by fusing together 100 parts of copper with 50 of grain tin.

V. Pewter.

Before the improvements in earthenware manufacture pewter was much used at table as plates, &c. At present it is employed to make dish-covers, measures for malt liquors, and teapots.

Pewter used to consist of lead and tin, in the proportion of one of the former to four of the latter. Too much lead was often added, which, becoming oxidated, produced salts having poisonous effects. But a great part of the pewter now made for beer measures consists of 100 parts of tin, 8 of antimony, 2 of bismuth, and 2 of copper.

The variety of pewter used for making Britannia metal teapots is said to be composed of equal parts of tin, antimony, bismuth, and brass; but the proportions probably vary in every manufactory.

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VI. SOLDER.

The process of soldering consists in uniting the surfaces of metals by the intervention of a more fusible metal. This is melted upon each surface; and, partly by cohesive force, partly by its melting the surface of the metal and forming an alloy, makes the two cohere.

Different alloys are used for soldering different metals. Tin plates are soldered together with an alloy consisting of from 142 parts of tin to 1 of lead; solder for pewter contains in addition bismuth; silver is usually soldered with pure tin, gold with an alloy of gold, and silver and platinum with pure gold.

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VII. ALLOYS FOR COINS.

Gold and silver are so soft that, were coins constructed solely of them, they would rapidly wear away. They are therefore always alloyed with copper, in the proportion of about one part in twelve.

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VIII. German or Nickel Silver.

This is composed of 25 parts of nickel, 25 of zinc, and 50 of copper; but the composition doubtless varies in different manufactories.

The Chinese pack-fong is a similar alloy. It is said to consist of 5 parts of copper alloyed with 7 of zinc and as many of nickel.

* * *

IX. Fusible Metal.

This is composed of 8 parts of bismuth, 5 of lead, and 3 of tin. It fuses at 202°, and serves to take casts of anatomical preparations.

* * *

X. Type Metal.

This consists of 3 parts of lead and 1 of antimony. That used for taking stereotype plates contains twice as much lead.

* * *

XI. Amalgams.

The amalgam employed in making mirrors consists of 1 part of lead, 1 of tin, 2 of bismuth, and 4 of mercury; 1 part of zinc, 1 of tin, and 2 of mercury constitute the amalgam used in the electrical machine; and an amalgam used for gilding brass is composed of 1 part of gold and 8 of mercury. But the mercury of this is afterwards expelled by heat.

CHAP. IX.

GENERAL PROPERTIES OF SALTS.

HITHERTO we have considered the properties of the elementary bodies and of some of those binary compounds which by virtue of their various affinities they form with one another. As before stated, these chemical affinities are supposed to be in some way dependent upon the electric condition of the elements, and two elements that combine together to make a compound are believed to do so in consequence of having opposite electric energies. This belief has affected the nomenclature, and the name that expresses the genus to which a binary compound belongs is always derived from the electro-negative element. Thus, in naming the compound formed by the union of one atom of nitrogen and one of oxygen, we do not use the expression nitruret of oxygen, but oxide of nitrogen; because oxygen has a greater electro-negative energy than nitrogen. So we say sulphuret of potassium not potassuret of sulphur, because potassium has a less electro-negative energy than sulphur; and so on for the rest.

None of these binary compounds can unite with an elementary body. Sulphuric acid, for example, can form no compound with metallic iron. Some organs of the binary compounds show no disposition to unite with other binary compounds; as, for instance, the peroxide of lead and other compounds noticed in Chap. VIII. But usually these binary compounds have very strong affinities for one another, do combine together, and form sundry compounds having altogether different properties from those of the two binary compounds the union of which makes them.

The most striking instances of this tendency of binary compounds to unite together is to be seen in the acids and alkalies. The affinities between these two classes of compounds is very strong, and they form by their union *salts*. Thus the common acid sulphuric acid, forms with the alkalies oxides of potassium, and oxide of soda the salts sulphate of potassa and sulphate of soda. In such instances the acids are endowed with the electro-negative energy (they generally contain a preponderance of oxygen), and the alkalies with the electro-positive. Hence the name of the genus of this class of compounds is indicated by the importance given to the electro-negative element sulphates, nitrates, phosphates, with the electro-positive alkali added.

Those salts formed by the union of two binary compounds, an acid and an alkali, often combine with water or with other salts to form still more complex compounds. To the distinct formation of a decided compound by the union of two salts, the name of double salt is given. In such instances of the two combining salts, the one has an electro-positive, and the other an electro-negative energy.

In the infancy of modern chemistry very exact definitions of acid, alkali, or base, and of the salts formed by the union of the two previous, were given; from all which, as the science has advanced, it has been found necessary to depart.

For instance, an acid was formerly considered to be characterised by three properties; 1st, that it contained oxygen; 2nd, that it had a sour taste and reddened litmus paper; and, 3rd, that it neutralised alkalies. The discovery of hydrochloric and other acids containing no oxygen, but acting in every other respect like acids, completely banished the first of these tests. Then some compounds which, from their insolubility, are neither sour nor characterised by reddening, were clearly determined to belong to this class of chemical compounds. And, lastly, some acids universally admitted as such, as the carbonic, were not quite able to destroy the alkalinity of potassa. Accordingly at present any substance is regarded as an acid which unites with potassa or other recognised alkalies, forming by such union bodies similar in their general chemical character to the compounds which sulphuric or other unquestioned acid forms with such alkalies.

A similar latitude is also now extended to alkalies or rather bases. The old definition of an alkali was causticity, an alkaline taste, and action upon colours, and the power of neutralising acids and thereby forming neutral salts. At present, although these properties are considered essential to an alkali proper, yet the last of them is the only one now considered essential to form a salifiable base; and any binary compound that unites with an admitted acid, as the nitric or sulphuric, and forms with such compounds analogous to those formed by the union of an acid and alkali, is regarded as a base.

In like manner, and of necessity, a very extended meaning began to be given to the expression *salt*. For long it was confined to the combination of an acid (containing oxygen) and a base (also containing oxygen). And these salts of which both acid and base contain oxygen, still constitute a very large and important class of salts called oxy-salts. But ever since the discovery of hydrochloric and other acids that contain hydrogen instead of oxygen, but which form with bases salts exactly analogous to the oxy-salts, these compounds have long been considered as true salts, and receive the appellation of hydro-salts.

Another class of salts is now universally admitted. The reader cannot have omitted to notice how closely sulphur corresponds in its chemical relations with oxygen, and he will not be surprised to learn that as there are oxy-salts,—saline compounds, in which the acid and the base both contain oxygen—so there are sulphur-salts, in which both the electro-positive and the electro-negative binary compounds, the union of which forms them, contain sulphur. Among the sulphur bases may be mentioned the protosulphurets of potassium, sodium, magnesium, &c., and among the sulphur acids sulphurets of arsenic, antimony, &c.

There is yet another class of salts, first distinguished as a class by Berzelius, and now generally admitted. These are called haloid salts (from *ἅλς*, *seawater*, and *εἶδος*, *form*), because in constitution they resemble sea salt, or chloride of sodium; that is to say, they are not formed by the union of two compound bodies, but are binary compounds of two elementary ones. According to this view all the chlorides, iodides, bromides, and fluorides are haloid. Whether these compounds approximate more to salts than to the oxides and sulphurets it is not necessary here to consider. In an elementary work they certainly seem to be more conveniently considered along with the oxides, and as such are noticed in Chaps. VI. and VII.

But, as the oxysalts unite with one another, forming double salts, so do these haloid bodies; and as such have been described as haloid salts by some systematic writers. Their example may be conveniently followed, and hence we have in this chapter to notice four different classes of salts*:—

I. The oxysalts, both acid and base of which contain oxygen.

II. The hydrosalts †, both acid and base of which contain hydrogen.

III. The sulphur salts, both acid and base of which contain sulphur.

IV. The compound haloid salts, both acid and base of which contain haloid salts.

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[*] As will be afterwards noticed; ammoniacal salts are an exception.

[†] It is important to remember that salts such as were formerly called hydriodate of potassa, &c.,—a name that seemed to imply that they were compounds of a hydro acid and an oxy base,—are now believed to consist of a haloid salt, iodide of potassium, &c., and water.

Before proceeding to details, however, it will be proper to consider some general properties of salts.

Salts differ very much in their relations with regard to water. Some are deliquescent, and attract the water from the atmosphere, becoming liquid. Nitrate of lime is an example of this. Others only attract water from a damp atmosphere; while some, again, undergo no change whatever, even if exposed to an atmosphere loaded with watery vapour.

So also do salts vary very much with regard to their solubility in water. Some, indeed, are quite insoluble in that fluid; some require many hundred times their own weight to dissolve them; while others freely melt in their own weight of water. As mentioned in Chapter I., during solution there is usually great reduction of temperature.

As the water in which any salt is dissolved evaporates, particularly if the evaporation take place slowly, the salt is deposited in regular crystals. As the particular shape of the crystal of every salt is, under the same circumstances, constant, these crystals constitute an important physical character, by means of which salts can be distinguished one from another. A statement of the principal forms of crystals forms the conclusion of this chapter.

During the act of crystallising a great many salts combine with a definite quantity of water, called the water of crystallisation, and which is essential to the formation of the crystal. Some crystallised salts when exposed to the air part with a portion of their water of crystallisation, and crumble down as a powder. Such are said to effloresce. Also this water of crystallisation may in all cases be expelled by heat.

In some cases, however, a portion of the water is so forcibly retained by the salt that it cannot be expelled save by a temperature that decomposes the salt. In some of these instances the water appears to act as a base, as will be again brought under our notice when we have to consider the phosphates. But it is now generally believed that the water thus obstinately retained by a salt exists sometimes in a state neither as water of crystallisation nor as basic water. This has been named by Graham, who first pointed this out, "constitutional" water. The speculations of this chemist upon this water, and the constitution of salts in connection with it, have an important bearing upon the philosophy of chemistry.

Salts when they crystallise often mechanically enclose water within their structures. If such salts are heated the water expands, and the salt is burst into fragments. A crackling noise attends this process, which is called decrepitation.

By the action of the voltaic battery salts are decomposed.

It is necessary now to consider the various forms assumed by crystals.

A work treating on chemistry cannot explain at length the mechanical system of crystallography, which has now become a very complex affair. But as reference must be made, when noticing salts, to the forms that their crystals assume, and as few general readers are familiar with crystallography, the following brief outline is necessary.

Six systems of crystals exist, which may be grouped into three classes, as follow:—

CLASS I.	Monometric,	containing one system,		viz.	1. *The regular or cubic.*
CLASS II.	Dimetric	„ two „	Quadratic -	„	2. *The square prismatic.*
			Hexagoual -	„	3. *The rhombohedric.*
CLASS III.	Trimetric	„ three „	Erect - -	„	4. *Right prismatic.*
			Oblique - -	„	5. *Oblique prismatic.*
			Doubly oblique	„	6. *Doubly oblique prismatic.*

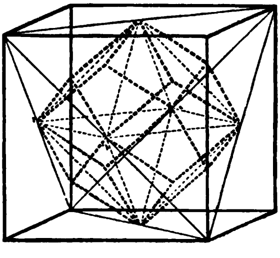
CLASS I. MONOMETRIC.

Essential character.—Three rectangular and equal axes. The refraction of light single, and the expansion by heat the same in all directions.

SYSTEM 1. REGULAR OR CUBIC.

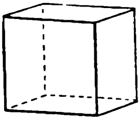
Essential character.—As there is only one system in the class, this is of course the same as that of the class. See *fig.* 49.

Fig. 49.



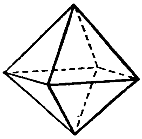
To this system belong

Fig. 50.



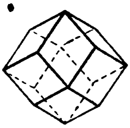
The Cube.

Fig. 51.



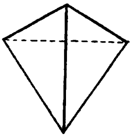
Regular Octohedron.

Fig. 52.



Dodecahedron.

Fig. 53.



Tetrahedron.

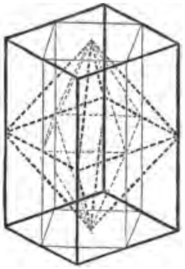
CLASS II. DIMETRIC.

Essential character. — The axes not rectangular and equal, but two of them equal; the refraction of light equal in all directions save one end; the expansion by heat also equal in two directions.

SYSTEM 2. SQUARE PRISMATIC.

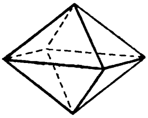
Essential character. — The axes three, rectangular and two equal, and belonging to CLASS II. See *fig.* 54.

Fig. 54.



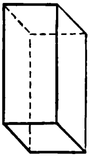
To this system belong

Fig. 55.



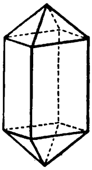
Octohedron with Square Base.

Fig. 56.



Right Square Prism.

Fig. 57.



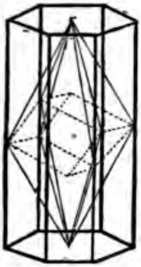
Combination of Octohedron and Prism.

In *fig.* 57. is shown the combination of the octohedron and prism.

SYSTEM 3. RHOMBOHEDRIC.

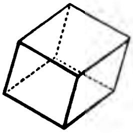
Essential character.—Axes four; three of them equal and placed in a plane; they cross at angles of 60°. The fourth axis differs from the others in length, and is perpendicular to them. See *fig.* 58.

Fig. 58.



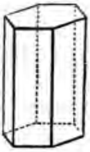
To this system belong

Fig. 59.



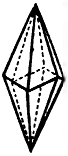
Rhombohedron or Rhomboid.

Fig. 60.



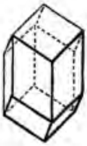
Hexagonal Prism.

Fig. 61.



The Scalenohedron.

Fig. 62.



Combination of Rhombohedron and Hexagonal Prism.

Fig. 62. shows a combination of the rhomboid with the hexagonal prism.

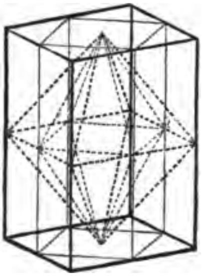
CLASS III. TRIMETRIC.

Essential character.—The axes three and unequal; the refraction of light double in all directions but two; and the expansion by heat unequal.

SYSTEM 4. RIGHT PRISMATIC.

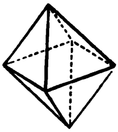
Essential character.—The three axes are rectangular. See *fig.* 63.

Fig. 63.



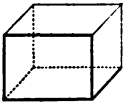
To this system belong

Fig. 64.



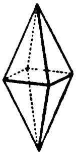
Octohedron with Rectangular Base.

Fig. 65.



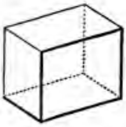
Right Rectangular Prism.

Fig. 66.



Octohedron with a Rhombic Base.

Fig. 67.

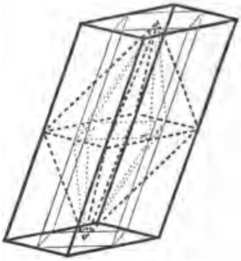


Right Rhombic Prism.

System 5. Oblique Prismatic.

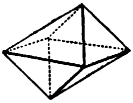
Essential character. — Axes three, and all unequal; and two of them cut one another obliquely, and are perpendicular to the third. See *fig.* 68.

Fig. 68.



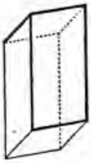
To this system belong

Fig. 69.



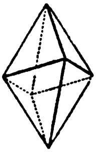
Oblique Octohedron with Rectangular Base.

Fig. 70.



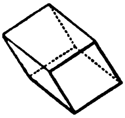
Oblique Rectangular Prism.

Fig. 71.



Oblique Octohedron with Rhombic Base.

Fig. 72.

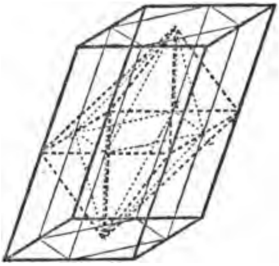


Oblique Rhombic Prism.

SYSTEM 6. DOUBLY OBLIQUE PRISMATIC.

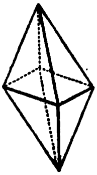
Essential character.—The three axes all oblique, and unequal, and angular to one another. See *fig.* 73.

Fig. 73.



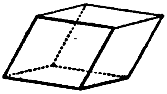
To this system belong

Fig. 74.



Doubly Oblique Octohedron.

Fig. 75.



Doubly Oblique Prism.

CHAP. X.

THE MORE IMPORTANT SALTS.

A. OXY-SALTS.

I. Sulphates.

All soluble sulphates may be readily detected by chloride of barium. A white powder, the sulphate of baryta, is at once thrown down, and is insoluble in acids or alkalies. A sulphate that is insoluble may be recognised by heating it in a platinum crucible, along with carbonate of potassa; mutual decomposition ensues, and if the residue be dissolved in water, and the potassa neutralised with hydrochloric acid, chloride of barium will throw down the characteristic white sulphate of baryta.

If a sulphate be heated along with carbonaceous matter the carbon combines with the oxygen of both acid and base, and a metallic sulphuret is formed.

1. Sulphate of Potassa.

This has received many names. The most common of its old-fashioned is *sal polychrest,* which literally means salt of many virtues. It may be prepared by adding sulphuric acid to carbonate of potassa; in which case the carbonic acid is expelled. It is usually obtained by neutralising with carbonate of potassa the residuum of the distillation of nitric acid (see p. 150.). It is saline and bitter to the taste, and is inodorous. It crystallises in single or double six-sided pyramids, the primary form of which is the right rhombic prism. The crystals contain no water of crystallisation, and are unacted upon by the air. They decrepitate when heated, and fuse at a red heat. They dissolve in six times their weight of boiling water, and sixteen of water at 60°. Sulphate of potassa is a remarkably hard salt.

2. BISULPHATE OF POTASSA.

This was formerly called acid, vitriolated tartar, and other names, and is now, as it contains an atom of basic water, called sulphate of potassa and water. It is prepared by adding sulphuric acid to the sulphate. It is crystallisable, the primary form of its crystal being a right rhombic prism. It has a very sour taste, and reddens litmus paper. It dissolves in its own weight of boiling water, and in twice its weight of water at 60°. Heated, it is resolved into sulphate and sulphuric acid.

3. SULPHATE OF SODA.

This compound, or Glauber's salts, occurs in animals and plants. It is obtained by adding sulphuric acid to carbonate of soda, and commercially it is procured in a large quantity from the residue for making chlorine. It has a saline, very bitter taste. It crystallises, the crystals commonly belonging to the right prismatic system. They contain ten atoms of water of crystallisation; the whole of which is lost by efflorescence on exposure to the air. One part of it dissolves in three parts of water at 60°, and in one at 212°.

4. Sulphate of Baryta.

This occurs native, and is called heavy spar. As such it is generally massive; but it is sometimes in crystals, the form of which is variable. It is easily formed by double decomposition, and is remarkable as being one of the most infusible and insoluble substances known.

5. Sulphate of Lime.

This is formed by adding sulphuric acid to any soluble salt of lime. It is found abundantly native. It occurs in the anhydrous state, in crystals belonging to the right prismatic system, and is called anhydrate. United with two atoms of water, one of which is held to be constitutional, it forms gypsum, the white compact variety of which is alabaster.

The water, however, of gypsum can be expelled by heat, and calcined gypsum constitutes plaster of Paris; which, when mixed with water, rapidly combines with it and becomes solid. As it so solidifies, it expands, and hence fills up the interstices of a mould in which it is put. On this account it is used for making casts.

It is a remarkable fact, however, that if the heat exceed 300° the sulphate of lime will not combine with water, but forms a permanent artificial anhydrate.

Sulphate of lime dissolves in about 500 parts of cold water, and, as it exists very abundantly in the earth, almost all spring water holds some of it in solution, and it is the common cause of *hardness*, as it is termed, in that fluid.

6. SULPHATE OF MAGNESIA.

This salt, commonly called Epsom salts, is present in many springs. For commercial purposes it is procured by the action of sulphuric acid upon carbonate of magnesia. It readily crystallises in small quadrangular prisms; but it may be obtained in larger ones, the principal form of which is a right rhombic prism. Sulphate of magnesia has a very bitter taste, is soluble in its own weight of water at 60°; its crystals contain water of crystallisation; heated to 212° they part with it save two atoms; one of these they lose at 270°, and the other at 460°.

7. SULPHATE OF ALUMINA.

By mixing sulphuric acid, water, and alumina a bisulphate is obtained which crystallises in thin flexible plates.

8. SULPHATE OF OXIDE OF MANGANESE.

There are two sulphates, one of the oxide and the other of the sesquioxide of manganese. The latter is prepared by dissolving the sesquioxide in sulphuric acid.

9. Sulphates of Oxides of Iron.

a. Sulphate of the Protoxide of Iron. This is very commonly called green vitriol, or green copperas. For commercial purposes it is prepared by the action of air upon the bisulphuret. In the laboratory it is obtained by the action of dilute sulphuric acid upon metallic iron. The iron is oxidated by the oxygen of the water, hydrogen is evolved (see p. 96.), and sulphate of the oxide formed.

Sulphate of the protoxide of iron crystallises in transparent green crystals, the primary form of which is an oblique rhombic prism. If exposed to the air, they absorb oxygen, and acquire, first a yellowish green tint, owing to their becoming converted into a sulphate of the black oxide, and then become brown, owing to a sulphate of the sesquioxide being formed. They are soluble in two parts of cold water; and the solution, if exposed to the air, becomes reddish, and deposits a sulphate of the sesquioxide. This, of course, is owing to the absorption of oxygen.

Sulphate of iron has an acrid, styptic taste. It crystallises with seven atoms of water, one of which is constitutional.

b. Sulphate of the Sesquioxide of Iron. This salt may be prepared by pouring sulphuric acid upon the red oxide.

c. Sulphate of the Peroxide of Iron. This may be formed by the action of nitric acid upon the protosulphate.

10. SULPHATE OF OXIDE OF ZINC.

This is sometimes known by the name of *white vitriol.* It is easily procured by the action of sulphuric acid upon water and metallic zinc; the decomposition that takes place being exactly analogous to that which happens in the preparation of sulphate of protoxide of iron. It is also made by roasting the native sulphuret of zinc. It crystallises in right rhombic prisms, which dissolve in two parts and a half of cold water and have a strong styptic taste.

11. SULPHATE OF OXIDE OF COPPER.

This is a sulphate of the protoxide, and is popularly named blue vitriol. It is best formed by adding diluted sulphuric acid to oxide of copper and then evaporating. It is deposited in fine blue crystals, the primary form of which is the doubly oblique prism. They contain five atoms of water; four of which are lost at a temperature of 212°, and the fifth at 430°. At this last temperature they fall down into a white powder, which, if brought into contact with water, combines with it, giving out considerable heat. Sulphate of protoxide of copper has a styptic, metallic taste, and dissolves in four parts of water at 40° and two of boiling water.

If ammonia be added to a solution of protoxide of copper, a suboxide is thrown down which is soluble in an excess of ammonia. This then forms a deep blue solution called ammoniated sulphate of copper, the exact nature of which is not very well known, but which is used as a test for arsenic. It is also prepared by rubbing sulphate of copper with sesquicarbonate of ammonia; in which case the carbonic acid of the latter is driven off.

12. SULPHATES OF OXIDES OF MERCURY.

If two parts of mercury are gently heated with three of sulphuric acid, and a sulphate of the protoxide is formed. If a strong heat be employed, the effervescence that takes place is more violent, more oxygen is derived from the air, and a bisulphate of the peroxide is produced. If this latter be thrown into hot water it is decomposed, free sulphuric acid separates and mixes with the water, and a yellow subsulphate, formerly called *turpeth mineral*, subsides.

DOUBLE SULPHATES.

13. SULPHATE OF LIME AND SODA.

This occurs native, constituting the glauberite of mineralogists. It may be artificially formed by fusing sulphate of soda and sulphate of lime together in the ratio of their atomic weights. It crystallises, forming oblique rhombic prisms.

14. SULPHATE OF POTASSA AND MAGNESIA.

If solutions of these two salts be mixed together in atomic proportions, and the common solution be evaporated, this double salt is obtained in crystals belonging to the oblique prismatic system.

15. SULPHATE OF POTASSA AND ALUMINA, OR COMMON ALUM.

The mode of preparing this important double salt varies. In this country it is usually obtained from aluminous shale; a mineral containing sulphur, iron, alumina, and, generally, potassa. It is extensively worked near Paisley and at Whitby; the shale at the latter place, however, being deficient in potassa, some salt of that substance requires to be added. The shale is roasted, by which the sulphuret of iron is oxidated, and the sulphuric acid combines with the potassa and alumina. By frequent crystallisation this double salt is detached from the oxide of iron.

Alum has an astringent and sweetish taste. It reddens litmus paper. It crystallises in octohedrons, and the crystals contain no less than 24 atoms of water. When heated they froth up very much and part with this water, becoming anhydrous alum, or the *burnt alum* of the shops. At a red heat they are decomposed, the acid being given off.

Alum is soluble in 18 times its own weight of water at 60°, and in its own weight of boiling water.

If alum is calcined with carbonaceous matter, a very inflammable substance, called Homberg's pyrophorus, is formed. This owes its properties to sulphuret of potassium, which is formed during the calcination.

Alum having exactly the same form, appearance, and taste as the salt just noticed, it may be formed by substituting ammonia or soda in place of the potassa.

Or, we can replace the alumina by another sesquibase. Thus, for instance, there is an—

a. Iron Alum. This is made by mixing sulphate of potassa with sulphate of peroxide of iron and evaporating. Crystals are obtained, in colour, form and taste, exactly the same as those of common alum.

b. Chromic Alum. This is prepared by the mixture of sulphates of potassa and chromium.

These alums present very good examples of isomorphism.

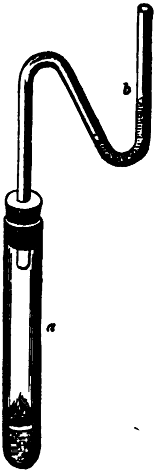
II. SULPHITES.

The salts formed by sulphurous acids are not important. They all have an alkaline reaction, and are very readily decomposed by other acids; and are also very apt to absorb oxygen and become converted into sulphates.

III. NITRATES.

These salts may be formed by the action of nitric acid, either on metals—in which case a portion of the acid oxidates the metal,—or on bases, or on carbonates. All the salts of nitric acid being soluble, it is impossible to precipitate this acid by any reagent. The nitrates, however, may be known by the following characters:—If heated with sulphuric acid, they disengage nitric acid, which may be recognised by its striking an orange-red colour with morphia; or the nitrate, copper filings, and sulphuric acid may be introduced into the tube *a*, *fig.* 76., to which a small curved glass tube *b* is attached. Heat is applied to *a*, nitric acid is separated; but the copper appropriates three of the atoms of oxygen, and binoxide of nitrogen is evolved. This binoxide blackens a solution of protoxide of iron. Accordingly a drop or two of solution of protosulphate of iron is put into the bend of *b*; and this, whenever the binoxide of nitrogen passes through it, becomes farther oxidated and of a black colour.

Fig. 76.



All the nitrates are decomposed by heat alone; and they are still more readily so decomposed when carbonaceous matter is mixed with them. Hence they are frequently used for oxidising other substances. The nitrate usually fixed upon for this purpose is the nitrate of potassa, or saltpetre. This is usually performed by mixing the inflammable body with an equal weight of the nitrate, and throwing the mixture in small quantities into a red-hot crucible. This process is called deflagration.

1. NITRATE OF POTASSA.

Nitrate of potassa is usually formed by putting decaying animal matter amongst soil which contains potassa. The nitrogen of the animal matter and a portion of the nitrogen of the air (as will be, it having a very important bearing upon farming, noticed in Book IV.) become oxidated and converted into nitric acid, which combines with the potassa. It naturally forms in hot countries in this manner, and is largely imported from the east. It is purified by repeated crystallisations.

Nitrate of potassa is a colourless salt, which crystallises in six-sided prisms. When pure, the crystals are transparent and colourless, and have a saline and cooling taste. It dissolves in its own weight of boiling water, but requires seven parts of water at 60°. During its solution in the latter a good deal of cold is generated; and hence it is used in frigorific powders. It has no water of crystallisation, although it usually contains some mechanically mixed with it. At 616° it fuses, and, like all the nitrates, at a red heat is decomposed.

It is used as an oxidating agent and in the manufacture of nitric acid; in making gunpowder, and frigorific mixtures; and, owing to its antiseptic properties, in the preservation of meat and butter.

2. Nitrate of Soda.

This is formed in hot countries by the action of the air upon salts containing some salt of soda. It may be artificially made by adding nitric acid to carbonate of soda. It crystallises in rhombs. It is very soluble in water, and deliquescent. It is employed in the manufacture of nitric and sulphuric acids, and as a manure.

3. NITRATE OF STRONTIA.

This may be made by digesting the native carbonate in nitric acid. Sometimes it crystallises in anhydrous octohedrons, and at others it contains a quantity of water of crystallisation; and then its crystals belong to the oblique prismatic system. It is used for making red fire at theatres, and for fireworks.

4. NITRATE OF PROTOXIDE OF COPPER.

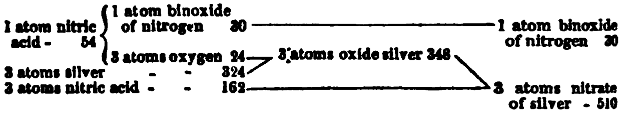
This salt is prepared by the action of dilute nitric acid on copper. It crystallises in oblique rhombs of a blue colour, which are very soluble and deliquescent. Heated up to 400°, it parts with a part of its acid, and a subnitrate remains.

5. NITRATES OF OXIDES OF MERCURY.

When mercury is dissolved in an excess of nitric acid with heat, a nitrate of the red oxide is formed. If this be dissolved in water, a basic nitrate, each atom containing three atoms of the oxide, is thrown down. If the mercury be dissolved in dilute nitric acid, without heat, a nitrate of the black oxide is formed.

6. NITRATE OF OXIDE OF SILVER.

This salt used to be called the infernal stone, and often yet receives the name of *lunar caustic.* It is prepared by mixing together silver and dilute acid, and then evaporating the water, The following table indicates the decomposition that occurs in this case and in similar ones: —


1 atom nitric acid - 54
1 atom binoxide of nitrogen 30
3 atoms oxygen 24
3 atoms silver - - 324
3 atoms nitric acid - - 162
3 atoms oxide silver 348
1 atom binoxide of nitrogen 30
3 atoms nitrate of silver - 510

Nitrate of oxide of silver forms transparent colourless crystals, the primary form of which is the right rhombic prism. Its taste is harsh and metallic. It fuses by heat; and when the temperature reaches 600° or 700°, decomposes, being resolved into oxygen, nitrous acid, and metallic silver. Exposed to the air it blackens, owing to the presence of any organic matter that may be in the atmosphere. The nitric acid would appear to attach itself to this; and the black colour is owing to the formation of an oxide of silver. Its action as a caustic is of the same nature. Nitrate of silver is soluble in its own weight of cold, and half its weight of hot, water. It is much used in the laboratory as a test for chlorine and hydrochloric acid.

* * *

IV. NITRITES.

The compounds of nitrous acid with bases have been little studied, and are of no importance.

* * *

V. CHLORATES.

The chlorates very much resemble the nitrates. They all consist of one atom of acid and one of base, are decomposed by a red heat, deflagrate with great violence with inflammable substances, and so readily evolve oxygen, that very slight causes produce in them violent explosions, and are all (with one exception, the chlorate of oxide of mercury) soluble in water.

Chlorates may be known by the action of sulphuric acid upon their solutions, which causes the evolution of the very peculiar smelling peroxide of chlorine.

1. CHLORATE OF POTASSA.

This was formerly known by the name of *oxymuriate of potash.* It is prepared by passing chlorine gas through a solution of potassa. The solution, after having been boiled for a few minutes, contains chloride of potassium and chlorate of potassa. If a portion of the water be evaporated, the chlorate of potassa crystallises, the chloride of potassium remaining in solution.

These crystals belong to the oblique rhombic prism system. They taste cool, dissolve in sixteen times their weight of cold water, and in two-and-a-half of boiling. Rubbed in the dark, they become luminous. Owing to the readiness with which it supplies oxygen, it is the basis of many fulminating mixtures, and, amongst other uses, smeared over a brimstone match it constitutes lucifers.

2. Perchlorate of Potassa.

This salt is only of importance owing to its being the least soluble salt of potassa. Hence perchloric acid is a test for potassa.

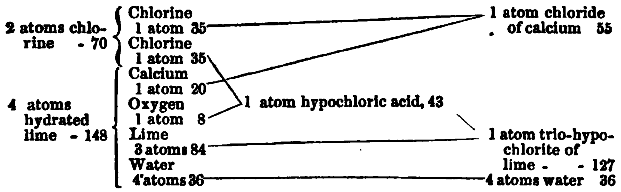
* * *

VI. Chlorites.

Chlorites of alkalies are easily made by transmitting a stream of chlorous acid gas through the solutions of pure alkalies. They are also possessed of very strong bleaching and oxidising properties. They may be known by the evolution of chlorous acid when sulphuric acid is put to them.

1. Hypochlorite of Lime.

Towards the close of the last century, Mr. Tennant, of Glasgow, took out a patent for a substance called Tennant's bleaching powder. Great difference has prevailed regarding the chemical nature of this substance. It is formed by passing chlorine through lime; and many suppose that it is a compound of chlorine and oxide of calcium, a chloride of lime. Others, with greater probability, think that one portion of the chlorine unites with so much of the calcium and forms chloride of calcium, while another portion of the chlorine unites with the oxygen to become hypochlorous acid, which combines with the undecomposed lime to form the compound in question. According to this view, Tennant's bleaching powder is a mixture of chloride of calcium and hypochloride of lime. The following diagram illustrates this: —



It is a dry, white powder, that smells of chlorine and has a strong taste. It dissolves partially in water, and the solution has very powerful bleaching properties. These depend upon the chlorine that the powder contains; and, since that found in commerce is not equally pure, its value is ascertained by the power of a given quantity to discharge a solution of indigo of known strength.

Hypochlorite of lime is also much used to destroy putrid odours, to check putrefaction, and, although this action is by no means so certain, to destroy the matter of contagion.

2. HYPOCHLORITE OF SODA.

This is prepared by passing chlorine through a solution of carbonate of soda. A decomposition analogous to the last takes place, and chloride of sodium, bicarbonate of soda, and hypochlorite of soda are formed. Its solution is called the disinfecting liquor of Labarraque.

VII. Iodates.

These salts closely resemble the chlorates, and, like them, form deflagrating mixtures with carbonaceous matter. Deoxiding agents, too, readily deprive the acid of these salts of their oxygen, and sulphurous acid, for example, sets the iodine free. This also serves as a test for them.

Iodate of Potassa.

This may be formed by adding iodine to a concentrated solution of potassa. Iodide of potassium and iodate of potassa are formed. The former is dissolved out by boiling alcohol, the iodate being insoluble in that menstruum. It is not a compound of much interest.

* * *

VIII. Phosphates.

According to the old view there are three isomeric phosphoric acids,—the phosphoric, the pyrophosphoric, and the metaphosphoric; each of which forms with bases salts. The characteristic of these acids is, that the first of them unites with one atom of a base, the second with two, and the third with three.

The prevalent opinion regarding their constitution now is, that the so-called phosphoric acids are in reality salts (viz., phosphates of the oxide of hydrogen, or water), and that—

Phosphoric acid is $(2P+5O)+(HO)$.
Pyrophosphoric acid is $(2P+5O)+2(HO)$.
Metaphosphoric acid is $(2P+5O)+3(HO)$.

Accordingly the first kind combines with one atom of an alkaline (or metallic) base to form a salt, the second with two, and the third with three. Hence they are now commonly called the monobasic, the bibasic, and the tribasic phosphoric acids.

Recently quite another view has been proposed; which is mentioned here not that it is necessary in an elementary book, but as indicative of a strong tendency that is growing among modern chemists to altogether give up the common opinion with regard to the constitution of acids and salts, and to consider the acid to be a compound of a radical* with hydrogen, and that the hydrogen may be replaced by an equivalent of metal, and that when this takes place a salt is formed. According to this manner of regarding the matter, the three phosphoric acids are all distinct ones, and are composed as follows:—

$2P+6O+H$.
$2P+7O+2H$.
$2P+8O+3H$.

As a hydrogen acid forms a neutral salt by the substitution of an atom of metal for each atom of hydrogen, these acids must necessarily require one, two, and three atoms of base to neutralise them, and must therefore produce different salts.

According to the older or commoner view the protophosphates are the compounds of the phosphoric or monobasic acids, and (and in agriculture this is turned to great practical advantage) they are all soluble in water. They also redden litmus paper, and hence in common language are often called superphosphates. The triphosphates (except those of the pure alkalies) are insoluble in water, but can by the action of strong acids be converted into the soluble phosphates.

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* A salt radical is that part of a salt which is non-metallic. The metal part is now often called a salt basyle.

A soluble phosphate is known from all other salts save an arsenite, by throwing down a yellow precipitate (phosphate of oxide of silver), when a solution of nitrate of oxide of silver is added to it. And it is known from an arsenite by not being changed in colour nor yielding a precipitate when a stream of sulphuretted hydrogen is passed through it.

1. TRIPHOSPHATE OF SODA (AND BASIC WATER—TURNER).

This is often named the rhombic phosphate, owing to the shape of its crystals. It is prepared by neutralising superphosphate of lime by carbonate of soda. Each atom of it is associated with no less than twenty-five of water. It has a very alkaline reaction.

2. TRIPHOSPHATE OF SODA AND AMMONIA (AND BASIC WATER).

Prepared by mixing together the above rhombic phosphate of soda with hydrochlorate of ammonia in a solution of hot water. It is much employed as a flux, and is often called microcosmic salt.

3. PHOSPHATE OF LIME.

This exists in bones, and is obtained by calcining them. It is not soluble in water.

4. TRIPHOSPHATE OF LIME.

This, the superphosphate of the farmers, is obtained, as above explained, by the action of sulphuric acid on the preceding.

IX. ARSENITES.

The arsenites of potassa and soda may be prepared by acting upon these alkalies by arsenious acid. They are very soluble in water, and, as stated under arsenious acid, form a yellow arsenite of oxide of silver, when mixed with the nitrate of that base; green arsenite of peroxide of copper when mixed with sulphate of that oxide; and yellow orpiment when mixed with sulphuretted hydrogen. Arsenite of potassa is used in medicine.

X. CHROMATES.

These salts are usually of either a yellow or a red colour, the latter tint being most developed when the acid is in excess. They may usually be distinguished by their colour, but any of them may also be recognised by boiling them in hydrochloric acid mixed with alcohol, upon which the chromic acid is first set free and decomposed, a green solution of chloride of chromium being produced.

1. Chromates of Potassa.

Bichromate of potassa is extensively manufactured. It is obtained by heating to redness the native oxide of chromium and iron along with nitrate of potassa, upon which chromic acid is generated, which unites with the potassa. A neutral chromate is obtained in a pure state by solution in water and crystallisation. To a solution of this sulphuric acid or acetic acid is 'added, which combines with some of the potassa, and the bichromate falls down in four-sided tabular crystals, that have a rich red colour.

2. Chromate of Oxide of Lead.

This is prepared by mixing together nitrate of lead and bichromate of potassa. Chromate of lead, or chrome yellow, is precipitated as a fine, lemon-yellow-coloured powder.

By adding a warm solution of potassa to the neutral chromate of lead half the acid is removed, and a dichromate of lead is formed, having a bright vermilion colour, and named chrome red.

Various mixtures of these two are made, having different shades of orange and red, and sold as pigments.

3. Chromate of Oxide of Silver.

This is prepared by mixing solutions of bichromate of potassa and nitrate of oxide of silver. If dissolved in nitric acid it is deposited in tabular crystals of a rich crimson colour.

4. CHROMATE OF OXIDE OF MERCURY.

This is formed by mixing a solution of chromate of potassa and subnitrate of mercury. It is of a brick-red colour.

* * *

XI. BORATES.

The boracic is so feeble an acid that the borates of soda and potassa have an alkaline reaction, and are readily decomposed by the stronger acids. They are all remarkably fusible in the fire.

Borates may be easily recognised. Any of them digested in sulphuric acid is decomposed; and if the solution be evaporated and the residue dissolved in alcohol, the alcoholic solution burns with a green flame.

BIBORATE OF SODA.

This occurs native. It crystallises in prisms belonging to the oblique system, which effloresce on exposure to the air, and dissolve in twenty parts of cold and six of warm water. Each atom is united with ten of water. This may be driven away by heat, and then the compound becomes a vitreous transparent solution,—glass of borax.

By a little management in the crystallisation it may be obtained in octohedrons, which are not efflorescent, and which only contain five atoms of water.

It is used as a flux, and the latter kind is thought the best suited for that purpose.

B. HYDRO-SALTS.

These as well as the members of the two other classes of salts only require a very short notice.

Hydro-salts are those the base and acid of which contain hydrogen. At one time the neutral compounds formed when hydrochloric, hydriodic, and other similar acids were added to oxidised bases were considered hydrochlorates, hydriodates, &c.; but it is now believed, as before stated, that the chlorine and iodine of the acid combine directly with the metal, forming a chloride, iodide, &c., and that the hydrogen of the acid and the oxygen of the base unite and form water.

The compounds of ammonia (a substance to be described in a subsequent Book) with the hydracids constitute the most important kind of hydro-salts; and even these, as will be subsequently mentioned, may be compounds of chlorine, &c., with the hypothetical basyle ammonium.

The only hydro-salt, then, that falls to be noticed here, and that more as an example than for any interest that it possesses, is —

HYDRIODATE OF PHOSPHURETTED HYDROGEN.

This consists of one atom of acid and one of base. It crystallises in cubes. The crystals are only permanent when dry, and if mixed with water yield a solution of hydriodic acid, and phosphuretted hydrogen gas, which escapes.

C. SULPHUR SALTS.

These compounds are double sulphurets, both base and acid containing sulphur, just as both base and acid of the oxy-salts contain oxygen. The more important sulphur acids are the protosulphurets of potassium, sodium, &c., and the sulphur bases arsenic, antimony, &c. Two examples of this class of salts will be sufficient.

1. HYDROSULPHURET OF SULPHURET OF POTASSIUM.

In this compound hydrosulphuric acid is the electro-negative ingredient, and sulphuret of potassium the positive one. Each atom of it contains an atom of its acid and of its base. It may be prepared by saturating a solution of pure potasssa with hydrosulphuric acid. At first the potassa exchanges elements with the gas and water, and protosulphuret of potassium are formed; and then the protosulphuret unites with the excess of hydrosulphuric acid. By careful evaporation, the double sulphuret is obtained in four or six-sided prisms.

2. Arsenio-Persulphuret of Sulphuret of Potassium.

This is a compound of sulphuret of arsenic with sulphuret of potassium. It is a salt of a lemon-yellow colour, in which rhomboidal tables may be distinguished.

D. HALOID SALTS.*

One or two examples of these will be quite sufficient.

1. Platino-bichloride of Chloride of Potassium.

This salt was formerly noticed as being produced when perchloride of platinum is added to potassa (see pp. 168. and 288.) It is usually obtained as a yellow powder, but if slowly evaporated it crystallises in octohedrons. If heated to redness, it yields chlorine, and the residue consists of platinum and chloride of potassium.

2. Oxychloride of Chloride of Iron.

This class of salts used to be called submuriates, on the supposition that they consist of hydrochloric acid combined with two or more atoms of an oxide. The one under notice is of a green colour, and is believed to be a compound of chloride of iron and protoxide of iron.

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* Double haloid salts of most writers.

Such is an account of the elementary bodies, and of the compounds which we witness them forming in the laboratory. Some of these compounds, indeed, they form in nature; but most of the above combinations either take place under the torture of fire or the arbitrary tyranny of the chemist. In nature their unions are more complex; and even in the dead world, while a body so simple as a binary compound is rare, an elementary body is scarcely ever seen alone and unassociated with other substances. We pass, then, from the simplification of the laboratory to consider how the elementary bodies, the properties and affinities of which we have been considering, are really arranged in the minerals and rocks that constitute the crust of the globe.

BOOK II.

THE CRUST OF THE EARTH,

OR THE

CHEMISTRY OF GEOLOGY.

CHAPTER I.

THE COMPOSITION OF THE MOST IMPORTANT OF THE MINERALS.

THE chemical elements of which the inorganic world is composed rarely exist separately. Usually several of them unite together; and when these compounds thus formed are solid, they are called minerals. The different minerals differ from one another, not only in chemical composition, but in their external and physical properties. Many of these are of very great importance, and now demand our consideration.

I. QUARTZ.

This mineral is the most abundant of any. It is present in almost every rock; and in most of the earlier ones it occurs in very large quantities. Unmixed with other minerals, and in a comminuted state, it forms the immense masses of matter that constitute the sands of the desert and of ocean beds.

Quartz is essentially oxide of silicon, and is composed of one atom of silicon and one of oxygen; but in nature it almost always has associated with it small quantities of alumina, lime, and iron, and sometimes magnesia and manganese, but these substances rarely amount to five per cent.

Quartz readily scratches glass and strikes fire with steel. Two pieces, rubbed together in the dark, become luminous. It is sometimes transparent, sometimes translucent, and sometimes opaque. When pure, it is colourless; but it occurs in such abundance, that it is mixed up with many other substances, and then becomes of various colours. Also, for the same reason, it has been subjected to different degrees of pressure; and it is to these two circumstances that is owing the many subvarieties of quartz that occur*; all of which, however, are essentially oxides of silicon. Among the more important of these subvarieties we may enumerate —

a. Flint. This is usually of a dark-grey colour, and consists of about 98 per cent. of oxide of silicon, with a little lime, alumina, and iron. It is found in the chalk formation, and is the quartz of the older rocks that have formed these, mechanically altered by the changes it has gone through. Flints, when exposed to the air, in time crumble down, and then constitute that variety of sand called gravel. Flints used to be much used for guns and for glass-making, but percussion caps and pure silicious sand are now generally substituted for them.

b. Rock Crystal. This is a nearly pure quartz, distinguished by being always transparent and generally white or brown. It is usually found in the primitive rocks. It is principally used as a gem, and made into seals, necklaces, &c., and, when large, into drinking-cups. These last were held in great esteem by the Romans. Sometimes rock crystal is of an orange colour, in which case it is often called a topaz, and in Scotland a cairngorum. Its form of crystallisation is a hexagonal prism.

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* Quartz is usually found in hexagonal crystals.

c. Amethyst. This is also a pure quartz, remarkable for its purplish-violet colour, and which occurs, although sparingly, in a great many formations.

d. Jasper. This is quartz with from nearly two to four per cent. of oxide of iron, which communicates a colour to it usually of a reddish-brown nature. It occurs in a great many formations, and is employed for ornamental purposes.

e. Chalcedony. This variety of quartz includes the carnelians (generally blood-red), and the heliotropes (commonly dark green with deep-red spots) or bloodstones.

f. Opal. This is a hydrated quartz, sometimes containing as much as ten per cent. of water. It is generally of a milk-white colour and semi-transparent. It is found in primitive and trap formations, and is used as a jewel.

g. Hornstone. This is a compact, translucent form of quartz, probably not very pure. It is common in many secondary and tertiary formations, and frequently contains petrifactions. This name hornstone is given to it from its somewhat resembling horn in colour, fracture, and transparency. Its colour, however, is sometimes yellow or green. It is used as an ornamental stone and for mill-stones.

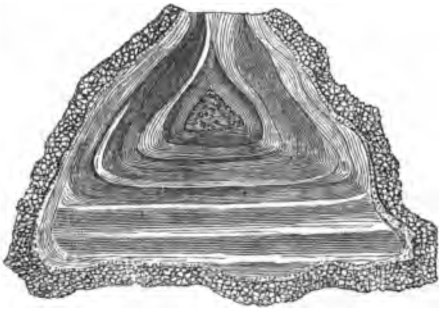
With regard to the above varieties of quartz we may observe, that the silicon and oxygen of which they are essentially composed probably united together the moment that these two elements were created, and that quartz has therefore existed as quartz ever since matter was called into being. And when we consider what has happened to almost every particle of quartz since that time,—how it has been made to change its position, exposed sometimes to intense heat, at others to extraordinary pressure, and at others again mixed with foreign substances,—we can easily understand how by this time the above varieties (and others not enumerated) have acquired their peculiarities. The formation of agate, the next variety to be noticed, although properly belonging to the next Chapter, is perhaps not quite so easy of explanation.

h. Agate. This is not a mineral, but a rock; that is to say, it consists of several minerals; but its chemical constitution is an oxide of silicon. It is in fact composed of a mixture of the above varieties, quartz, chalcedony, flint, hornstone, carnelian, jasper, and amethyst.

Agate, or Scotch pebble, as it is often called, occurs in masses usually two or three inches in diameter, but sometimes larger, imbedded in the amygdaloidal trap rocks. The agates have clearly been deposited in a hole formed in the trap, and also as clearly after the trap had cooled; and they are not cemented to the trap, but have commonly a thin layer of green earth interposed between their surface and that of the surrounding trap. The centre is very frequently a piece of quartz, and surrounding it are layers of the above-mentioned ingredients. Owing to these different layers a polished surface of agate presents a beautifully variegated appearance.

The trap rocks, as we shall soon have occasion to notice, are produced by igneous and volcanic action. As during eruptions of volcanoes much gas is evolved, we can easily understand that just when the trap was solidifying large bubbles of gas would be arrested in their transit upwards, and thus form the cavities in which the agates are deposited. It is more difficult to comprehend how so many varieties of quartz are formed in so limited a space. Perhaps the quartz flowed into this cavity, and the difference of the forms that it took depended upon—1st, various admixture with some of the amygdaloid, then not quite solid; and, 2dly, upon slight degrees of difference of intensity of heat and pressure. *Fig.* 77. illustrates the construction of the agate.

Fig. 77.



Agates are found pretty abundantly in the trap formations of Scotland, and in those of Germany, &c. They are used in the arts for making mortars in which to pound very hard substances, and also for making ornaments, snuff-boxes, and vases. About ninety-eight per cent. of agate is composed of oxide of silicon.

II. Felspar.

Quartz with its subvarieties is a compound of oxygen and silicon, and is indifferently called oxide of silicon or silicic acid. This silicic acid has a strong affinity for both alumina and potassa; and, as both these two latter substances occur very abundantly in nature, it is not surprising that during the many convulsions that have so frequently brought the different material elements of this world in approximation, a mineral has been largely formed by the union of silicic acid with these two bases. Such has been the case; and to this double silicate of alumina and potassa is given the name of felspar.

Common felspar is generally either white or red, sometimes grey, and more rarely blue or green. It takes upon it, however, a great many intermediate colours. It is generally massive and disseminated, but sometimes crystallised. It is hard, but less hard than quartz, heavy, and easily broken. It occurs so very abundantly, that the proportions of the ingredients of different specimens vary somewhat. Perhaps the following exhibits its average composition:—

Silicic acid - -	64 parts.
Alumina - -	20 ,,
Potassa - -	14 ,,
Lime - - -	2 ,,
	100 ,,

It also frequently contains a little soda, magnesia, and iron. Felspar is one of the most abundant minerals in nature; forming fully two-thirds of the older rocks, a considerable portion of igneous ones, and occurring also in other formations, generally associated with quartz and with hornblende and mica. When exposed to the air its chemical constitution is liable to be affected, and its elements enter into other arrangements immediately to be noticed. We proceed to enumerate some subvarieties of felspar. It usually occurs in rectangular prisms.

a. Compact Felspar. This is probably fused common felspar, with which a little fused quartz has combined. Sometimes it is harder than the common felspar; but it is frequently softer, and when it is so receives the name of clay-stone. It occurs in comparatively small quantities in the primitive mountains, but in considerable quantity in some of the transition and later formations. It is frequent, for instance, in the Pentland hills. A specimen of it from this last-mentioned source gave, when analysed, the following result:—

Silicic acid	71·17	
Alumina	13·60	
Potassa	3·19	
Lime	·40	
Oxide of iron	1·40	
Manganese	·10	
Volatile matter	3·50	probably water.
	93·36	
Loss	6·64	
	100·	

b. Albite. This is a variety of felspar in which some soda is substituted for part of the potash. Its appearances and physical properties are almost identical with those of common felspar, and its occurrence is only another illustration of the fact that soda and potassa are isomorphous substances. The following is an analysis of albite:—

Silicic acid - -	68·45
Alumina - - -	18·71
Soda - - -	11·24
Potassa - - -	·65
Lime - - -	·50
Magnesia - -	·18
Peroxide of iron -	·27
	100·

Albite is found, but in no very great quantities, in granite and the other more primitive rocks, but is more abundant in the igneous greenstones.

c. Obsidian. This is a brittle, compact mineral, or rather compound mineral; generally, but not always, of a velvet-black colour and of a vitreous aspect that indicates that it has been formed by the rapid cooling of heated masses, probably of felspar and (in a much smaller proportion) of opal. It occurs in volcanic countries, as Iceland and many parts adjoining the Mediterranean, but in far greater quantity in South America. It was formerly used in this last-mentioned continent for making mirrors and cutting implements. Cortes, in his letter to Charles V. says, that he has seen razors made of it. It is also found in New Zealand. The following is Murdoch's analysis of a specimen found in the Bay of Islands:—

Silicic acid - - -	75·20
Alumina - - -	6·86
Soda and potassa - -	7·57
Lime and magnesia - -	3·83
Peroxide of iron - -	6·54
	100·

d. Pumice. This is probably obsidian suddenly cooled when minute bubbles of gaseous matter were passing through it. At any rate it is found in the same localities, and has the same chemical constitution; but, owing to its mechanical condition being then affected, its specific gravity is much reduced, to such an extent indeed, that it swims upon water. It is used in this country for polishing mirrors, soft stones, leather, &c.; in the East, in the bath rooms, to remove superfluous hairs, and, according to Jameson, by the Mediterranean sailors to rub off their beards, as a substitute for shaving.

e. Kaolin. This is nothing else than felspar which has split to very minute particles by frosts and become permeated by water, one portion of which has been chemically combined with it, but another part of which has washed out some of the silicate of potassa. It is found in small quantities lying upon rocks containing felspar, such as granite. It is used for making the finest kinds of porcelain. It is distinguished from common clay by the extreme fineness of its particles, and by its soiling strongly. The following is an analysis of a specimen of it found in Devonshire:—

Silicic acid	47·20
Alumina with traces of iron and manganese	38·80
Magnesia and lime	·24
Alkali and loss	1·76
Water	·12
	88·12

f. Clay. Clay is simply a coarser kaolin, and is formed by the decomposition of felspar.* The basis of clay is essentially alumina, and some of the clays are perhaps nearly pure alumina, the silicates having been washed away. But clay almost invariably contains silicic acid, lime, potash, soda, &c. The appearances, &c., of clay are so familiar, that it is unnecessary to describe them. Two varieties, not however common, of clays are remarkable for containing a large amount of magnesia; these are soapstone and the meerchaum clay used for making pipe-heads. Very probably a great part of our agricultural clays likewise contain magnesia, and perhaps a little sulphur, phosphorus, &c.

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* Of course we mean the greater part of the clay has this origin. The minerals that contain alumina in small quantities may have produced a little of it.

It will be observed that the prominent character of the quartzes consists in their affording silica, while that of the felspars lies in the alumina which they contain.

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III. HORNBLENDE.

Hornblende is a mineral the colour of which passes from a greenish black to an olive-green, and sometimes to a liver-brown. The blacker-coloured hornblendes are opaque, but the edges of those having a greener hue are translucent. It is brittle, but not very hard. It is a very common mineral, forming nearly the entire mass of hornblende slate; united with felspar it composes syenite, and with albite greenstone rocks. It is also common in granite, gneiss, mica, and chlorite slates. When exposed to the air it is remarkably prone to crumble down; and no rocks perhaps so rapidly form soils as those that
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are rich in this substance. It occurs commonly in oblique, four-sided prisms. The following is an analysis of a specimen of it:—

Silicic acid	60·10
Magnesia	24·31
Lime	12·73
Iron (protoxide)	1·00
Manganese (do.)	·47
Alumina	·42
Fluoric acid	·83
Water	·15
	100·01

The most important subvariety of hornblende is—

a. Augite. The colour of this mineral is usually a greenish black. It is brittle, and has a higher specific gravity than hornblende. Like hornblende it is a constituent of many igneous rocks, but not the same igneous rocks as those in which hornblende is abundant. Augite occurs in basalt, clinkstones, &c. The following is an analysis of it:—

Silicic acid	54·15
Lime	24·74
Magnesia	18·22
Protoxide of iron	2·51
Protoxide of manganese	·18
Alumina	·20
	100·

When hornblende and augite are crumbled down and form soil, the magnesia they contain will doubtless be very valuable; but in the soil formed over primitive and igneous rocks their main use is dependent upon the lime which forms so large a portion of them. The lime obtained by the soil lying over later formations is, as we shall soon have occasion to notice, derived from another source.

We may remark that augite (as is also the case with felspar) is sometimes artificially formed in the sides of the furnaces of iron-smelting works.

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IV. Mica.

There are three varieties of mica proper, and also some other minerals so closely allied to it that we may put them down as subvarieties. Mica usually occurs in rhomboidal, four-sided prisms, or in regular six-sided ones. Each of these may be enumerated separately.

a. Potash-mica. Mica is remarkable for its splitting into laminæ. According to Barry a lamina of it may be obtained the 250,000th of an inch in thickness. This form of mica is very abundant. It is an essential constituent of granite, gneiss, and mica slates, and it is often present in a great many other formations. The analyses that have been made of it vary; owing probably, in a great measure, to its being mixed up with specimens of magnesia-mica. The following is an analysis made by Rose:—

Silicic acid - - -	47·50
Alumina - - -	37·20
Iron (peroxide) - -	3·20
Manganese (do.) - -	·90
Potassa - - -	9·60
Fluoric acid - - -	5·56
Water - - - -	2·63
	101·59

Rose always found fluoric acid, but the quantity in the above must have been accidental.

Potash-mica is sometimes formed artificially during metallurgic processes, when wood is the fuel employed. In such a case the potash is of course derived from the wood ashes.

b. Lithia-mica. This in physical appearance resembles the foregoing, save that it is often of a rose-red colour. It is found in granite, gneiss, &c., and is sometimes used as an ornamental stone. The following is one of the analyses made by Turner:—

Silicic acid - - -	50·91
Alumina - - -	28·17
Manganese (oxide) - -	1·08
Potassa - - -	9·50
Lithia - - -	5·67
Fluoric acid - - -	4·11
	99·44

It is said to contain also a small proportion of phosphorus in the form of phosphoric acid.

c. Magnesia-mica. The peculiarity of the composition of this is that it contains about twenty per cent. of magnesia and less alumina than the two foregoing.

d. Chlorite. This is a very important and widely-distributed mineral. It occurs in thin plates, which are translucent and sometimes transparent, and is remarkable for its green colour. It is found in granite, gneiss, and other primitive rocks, also in basalt and other igneous rocks, and is most abundant in chlorite slate. The analyses of it are variable. One gives the following results:—

Silicic acid - - -	25·51
Alumina - -	21·91
Magnesia - - -	22·83
Iron (protoxide) - -	15·
Water - - -	12·
	98·15

The *green earth* found in amygdaloid is probably decayed chlorite. The agates found in the amygdaloidal cavities, noticed above, are generally surrounded with a little of this substance.

e. Talc. This mineral resembles mica, but is distinguished from it by its very greasy feel, its softness, and by its colour, which is generally greenish. The mineral sometimes called steatite is said to be only a different form of talc.

Talc and steatite occur in many crystalline rocks, and are sometimes put to strange uses. Both are employed in the manufacture of rouge, and rouge made from them is said not to make the skin harsh. In the East they are employed for rubbing the body to make the surface smooth. The Chinamen mix the ashes of talc with wine, and use it as an *elixir vitæ;* but the most extraordinary employment of it is its being eaten by Indians and Africans to appease the pangs of hunger. Talc is composed of about sixty per cent. of silicic acid, and of about thirty of magnesia.

f. Serpentine. This mineral, usually of a green colour, is translucent at the edges, occurs massive, is soft, and feels somewhat greasy. It is found in primitive rocks (where the beds of it usually contain iron and copper ores) and in igneous rocks. As it is easily worked and takes a good polish, it is employed sometimes for making ornamental boxes, &c. Its chemical composition very much approximates to that of talc.

V. CALC SPAR.

Calc spar, or limestone, is distinguished by its effervescence when acids are applied to it, and by its complete solubility in nitric and hydrochloric acids. It is a carbonate of lime, and is composed entirely of the shells of dead animals. It of course occurs in every geological formation upon which animal life has had a being; that is to say, it is present upon the Silurian and all subsequent formations, and it is at this moment being largely deposited in the South Seas by the corals, myriads of whom are daily dying and leaving their osseous structures to form coral reefs or limestone deposits. As in the primitive world hornblende was the main source of lime to the soil, so now—and such has been the case for an immense period—plants have to obtain their supply of this necessary constituent mainly from this calc spar.

As limestone occurs so very abundantly, as it has therefore been exposed to very different degrees of pressure, and as it is necessarily occasionally intermixed with other substances, many kinds of it differ in their structure, colour, &c., to which distinctive names have been given. Thus, when it is compact and takes a fine polish, it is termed marble; and again of marble there are very many sub-kinds. There is a limestone in which the putrifying remains of the flesh, &c., of the animals are present, and which gives an unpleasant smell to it; whence it is called stinkstone: chalk is a crumbled-down sort of limestone; and when chalk is mixed with alumina or clay it is called a marl. An important subvariety of limestone is—

a. Dolomite. This mineral, which is sometimes named magnesian limestone, contains, besides carbonate of lime, from twenty-five to fifty per cent. of carbonate of magnesia, and usually also a small quantity of carbonate of iron. This magnesian limestone forms large tracts, especially in the north of England, and occurs in smaller quantities in many localities.

VI. Apatite.

This important mineral is found in almost all formations, although never very abundantly. It occurs in the primitive rocks, in mica slate, in igneous rocks, and probably into every one of the transition and subsequent formations. The following is an analysis of it:—

Lime	55·30
Hydrochloric acid	·43
Phosphoric and fluoric acids, and loss	44·27
	100·

The quantity of fluoric acid is small. In fact, the value of apatite depends upon its affording phosphorus to the soil; and we have reason to believe that all the phosphorus that exists in animals and plants has been originally derived from apatite, and indeed that all the phosphorus that will exist in future animals and plants must be derived either from dead bones, &c., or from this mineral.

It usually occurs in six-sided prisms.

VII. GYPSUM.

Gypsum is a mineral common to almost all the formations, but more abundant in the later ones. It is a sulphate of lime with twenty-one per cent. of water. It has unquestionably been formed by some of the sulphur originally created in this world becoming converted by contact with oxygen into sulphuric acid, which acid has subsequently gradually combined with the lime of the earlier primitive rocks.

The sulphur essential to all fertile soil is mainly derived from gypsum. There are, however, other sources of sulphur to soil; for sulphur itself occurs pretty extensively native, and some of the metallic sulphurets exist to a great amount.

VIII. ROCK SALT.

This very familiar mineral is most extensively distributed. In fact, if it does not exist in any part of the subsoil of a country, it is almost invariably present in the air over it, having been originally deposited there by evaporation from the sea, and conveyed from place to place by means of winds. Rock salt is composed of sodium and chlorine, and the greater part of these two elements that soil must possess is obtained from common salt.*

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* Another common mineral is *fluor spar*. It is fluoride of calcium.

The above eight minerals are of importance, because they are those that originally formed and still keep up the constantly wasting soil, and therefore it is they that form the greater part of the solid structure of plants, of animals, and of man. Regarding them in this light, they are of the greatest consequence to us, inasmuch as our material existence entirely depends upon them; and were one of them to become exhausted, all animal and vegetable life would (as will be soon explained) become quite impossible. Man, however, is not content with mere existence. Once, as we now know, he was ignorant of metals and the art of working them; and, mere savages, our ancestors dwelt in subterranean holes, their only tools for chase or industry being composed of stone.* Increasing intelligence discovered the metals; and now their constant employment is essential for our comfort, if not for our very being. Hence the minerals that contain metals, or the more important metallic ores, are of the greatest interest to us, and some of them demand at least an enumeration. Farther there comes a time in the history of civilisation when mankind, having attained the useful and indispensable, seek the gorgeous and the rare. Hence it will be necessary to run through the minerals that contain, or that are, gems. But now we have arrived at an epoch when industry and applied skill and science seek to enlarge the bounds of human comfort, perhaps the destinies of humanity. To do this we have pressed into our service certain other minerals; and these carbonaceous minerals, or coals, have so well played their part, that we cannot pass them over in silence. We proceed, then, briefly to notice the more important metallic ores, the gems, and the coals, all of which form part of the structure of the crust of our globe.

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* See Wilson's Præ-historic Annals of Scotland.

Metals generally occur combined with other substances. A few, however, are found alone, or, to use the technical expression, native. Two or three of these claim attention, and we begin with—

IX. NATIVE MERCURY.

Mercury occurs native, although very rarely, in all geological formations. Owing to its fluid state it is generally found at the bottom of a fissure. It is most abundant in Carniola and Spain, but is found in many other places, including, it is said, among others, the neighbourhood of Berwick.

X. NATIVE GOLD.

Gold is found native; sometimes, but rarely, in beds; but more frequently disseminated, particularly among quartz, and in grains or larger masses among sand. It is, however, entirely confined to primitive or metamorphic formations. As the more recent formations have been formed from the disintegration of previous primitive ores that probably likewise contained gold, it becomes a curious question, what has become of it? Perhaps it exists in the more recent formations in a state of such minute division as to have escaped observation.

Native gold has been found in the alluvial soil near Leadmills in Lanarkshire; a locality indeed that was once successfully worked as a gold mine. It has also been found amongst the granites of Cornwall and Wicklow; and indeed the largest "nugget" met with in Europe came from this last-mentioned locality. Several gold mines are still in operation in many parts of the old world and in South America, and many have been given up as unprofitable. In recent times most extensive gold fields have been discovered (or rediscovered) in California and Australia. It may perhaps be doubted, however, if these discoveries will have the wonderful effects that are sometimes anticipated.

XI. NATIVE IRON.

Small portions of native iron have occasionally been found, but we here notice it for the sake of mentioning meteoric iron, which is very interesting, inasmuch as it would seem to indicate that, in some respects at least, the chemical composition of the moon resembles that of our globe. For a long time it has been an object of popular belief that stones fell from the air upon the earth, and one of these stones was worshipped for long in a Syrian temple. But it is only within the last century that the evidence regarding them has become so strong that scientific men have recognised their existence. Many have fallen at many different parts of the world; but when analysed they all agree in affording a large quantity (about ninety per cent.) of native iron, and some of the very rare metal nickel, with varying traces of sulphur, cobalt, &c.

The most probable origin of these meteoric stones is, that they have been ejected with such force from some of the many volcanoes that exist in the moon, that they have fallen out of the attraction of that satellite into that of our planet.

XII. CINNABAR.

This mineral is the principal ore of mercury. It is found in the primitive and the igneous rocks, in the transition formations, and pretty abundantly in sandstone lying over coal. It is the sulphuret of mercury, and is extensively prepared by mixing sulphur and mercury for the purpose of forming a pigment.

XIII. COPPER PYRITES.

This mineral is the most abundant ore of copper. It is the sulphuret of that metal mixed with a little sulphuret of iron. It is found in a great many rocks, and occurs in granite, gneiss, mica-slate, clay-slate, igneous rocks, the graywacke, and the limestone, coal, and sandstone formations.

XIV. HÆMATITE.

This is one of several important ores of iron, and is an oxide of that metal. These ferruginous ores, along with other important economical metallic ores, will be again noticed. We proceed to name some of the more important minerals that are used for ornamental purposes.

XV. THE DIAMOND.

This most precious and esteemed of gems is a simple body, being composed only of carbon. The usual places where it has been found is over the most recent formations; and the diamond is usually considered to have a vegetable origin. But it would now appear certain that diamonds have been discovered in mica-slate; and as, during the time the mica slate was being deposited, there was no vegetation, many maintain that this form of carbon has some other origin than a vegetable one. The speculations that have been made upon the subject are not, however, very definite ones.

The most usual form of the diamond is in octohedrons.

XVI. HYACINTH.

This gem is often sold as a diamond. Its chemical constitution is, however, very different. It is a silicate of the rare earthy oxide zirconia. The following is an analysis of it:—

Silicic acid - - -	26·5
Zirconia - - -	69·0
Oxide of iron - -	·5
	96·0

It is principally found in the igneous rocks and in syenite. Its usual form is four-sided prisms.

XVII. SAPPHIRE.

This is next to diamond the most esteemed of the precious stones. It is principally found in alluvial soil formed by the decomposition of the more recent igneous rocks. It consists essentially of alumina, with a little lime and oxide of iron.

XVIII. EMERALD.

This gem when it does not possess a brilliant green colour is named a beryl. It contains two rare elements. The following is an analysis:—

Silicic acid - - -	64·40
Alumina - - -	14·00
Glucina - - -	13·00
Lime - - - -	2·56
Chrome oxide - -	3·50
	97·46

Tantalic acid has also been detected in some specimens. It occurs generally in six-sided prisms.

XIX. Coal.

Coal is composed of wood that has been subjected to strong pressure, and of course the chemical composition of the coal is the same as that of the trees that have furnished it. As we shall see in the next Book, all wood contains a very large quantity of carbon, and accordingly we find this element predominate in coal. Wood, too, as we shall likewise see, also contains oxygen, hydrogen, often nitrogen, and a small proportion of mineral matter known by the name of ash. As a matter of course coal contains these also. In whatever formation there was vegetation in that there may be coal; but as, during what is called the carboniferous period (soon to be noticed) the vegetation was characterised by extreme luxuriance, by far the greater quantity of coal is found in it. Of the varieties of coal we may enumerate—

a. Graphite. This substance, or blacklead as it is sometimes very improperly (for it contains no lead at all) called, is sometimes a coal, *i. e.* of vegetable origin. Thus, there is a coalfield in Ayrshire upon a portion of which has been found some recent volcanic rock, the heat of which has converted the coal near to it into graphite. The graphite used for making pencils in Cumberland is probably derived from the scanty vegetation of the Cambrian formation. But we sometimes find graphite in the gneiss; and in such an instance, the same as when we find diamond in the mica, we can only say that carbon in the very earlier days of the world has had a career of which we have now no knowledge.

Graphite is rich in carbon, and contains a small ash. The other substances, as iron, alum, &c., that have been found associated with it, are probably accidently so, although the iron is very constant in its presence, and graphite is often artificially formed in iron furnaces.

b. Anthracite. This is common coal altered by heat, and generally containing less oxygen and hydrogen. It is common in many of our coal fields, and is particularly so in some of the United States of America, where it is extensively used for economic purposes.

c. Common Coal. This is the usual coal formed out of the forests of the carboniferous era. The following is an analysis of a Welsh coal:—

Carbon - - -	85·52
Hydrogen - -	3·72
Sulphur - - -	·12
Oxygen - - -	4·55
Ash - - -	6·09
	100·00

d. Brown Coal. This is a more recently formed coal than the last, and is often only partially converted into coal. It is found in the oolite, the chalk, and the tertiary formation. It contains a much less proportion of carbon than common coal, and much more oxygen, hydrogen, and combined water, and therefore burns much worse. It is generally beat in troughs, made up into bricks, and dried before being used.

Of the above minerals the first eight, with their subvarieties, occur by far the most abundantly, and the student ought to be familiar with their chemical compositions. The minerals and the mineral ores are much more sparingly distributed among the crust of the earth, yet still in sufficient abundance to minister to the wants of industrious man. The gems, together with many other minerals that we have not enumerated, are rare, and for our present purposes in this work of very little consequence. Still it was necessary to notice them, in order to obtain a just notion of the manner in which the elements of matter are grouped together and arranged in the inorganic world of nature.

CHAP. II.

THE CHEMISTRY OF THE ROCKS.

We have already seen that all the objects that form the crust of this globe are composed of a comparatively few elementary bodies. Sometimes these elements occur alone and unassociated with other substances; and we can always isolate them in the laboratory. More frequently they occur combined according to certain fixed laws with other elements, forming compound bodies. Sometimes we find these compounds pure, and by art we can always make them so; but in nature it is usual to find even two or three compounds associated together, forming a homogeneous structure which has a texture, mode of breaking, &c., which distinguish it from all other substances.

Such are nearly all the minerals noticed in the preceding Chapter. Quartz, for example, is composed of oxide of silicon mixed with a little oxide of aluminum; felspar is a mixture of silicate of potassa, tersilicate of alumina, and a little lime; albite is a mixture of silicate of soda, tersilicate of alumina, a little potassa, lime, magnesia, and iron; and so on for the rest. Now, as a compound body is formed by the union of its elements, as a mineral is formed by the union of compound bodies, so is a rock composed by the union of minerals.

Geologically speaking, a rock is a mineral mass of sufficient size as to be regarded as an essential portion of the surface of the globe. It does not signify as in common language, an exposed, perpendicular, and hard surface; for it may be covered by other rocks, or lie in a flat position, and be either hard or soft.

It is not, however, quite correct to say that a rock is *always* composed of more than one mineral; for some minerals, as quartz, chalk, and others, do sometimes exist alone in such large masses as to constitute undoubted rocks. The following is a list of the more important of the rocks, and of the minerals of which they are composed; and by referring to the previous Chapter the reader may ascertain what the chemical composition of any rock is, and, therefore, as will be afterwards explained, what the chemical composition of the soil formed by its decomposition will be.

It is, however, proper to state here, that our knowledge regarding rocks is by no means so accurate and precise as it could be wished to be; and many rocks undoubtedly contain minerals which are not as yet set down in the books.

Of the rocks in which quartz is a prominent ingredient we may select—

Sandstone. This is nothing but quartz which has been water-carried and then compressed. Hence it is neither so hard, so compact, nor so close-grained as quartz in its more primeval state. It is often of a reddish colour, which is generally owing to the presence of iron. It, perhaps, also always contains apatite and minerals made up of lime, soda, chlorine, &c.

Among the rocks characterised by containing felspar the most important are —

a. Granite. This essentially consists of quartz, mica, and felspar, the last being the most abundant. But granite almost invariably also contains, although in less quantities, apatite and many of the other minerals mentioned in the last Chapter.

b. Syenite. This merely differs from granite in containing hornblende in place of mica. In both granite and syenite there are often distinct crystals of felspar imbedded in the granular mass, in which case both are said to be porphyritic. As a producer of soil, syenite mainly differs from granite in affording more lime.

c. Compact Felspar is principally a mixture of felspar and quartz.

d. Protogene. This rock differs from granite in having talc, steatite, or chloride in place of mica.

These three rocks are unstratified, and each has clearly been a fluid mass rapidly cooled.

e. Gneiss. This is merely a variety of granite which has been metamorphosed; *i. e.* derived from extinct granite mountains, and which, as it has been deposited by degrees, occurs in layers.

Among the rocks characterised by containing *mica* as the principal ingredient may be enumerated—

a. Mica-slate. This is a stratified rock, of which the two principal (it, doubtless, in general contains some of almost all those enumerated in the last chapter) minerals are mica and quartz. It is either metamorphic granite or gneiss.

b. Clay-slate. This is often probably metamorphic mica-slate. During its deposition the particles of mica have been very much rubbed and ground down, and hence it has a glistening, homogeneous appearance.

We may next notice the talcose rocks, as they are sometimes called.

a. Chlorite-slate. The two most important elements of this are chlorite and quartz.

b. Talc-slate. The two most important elements of this are talc and quartz.

Both these are very slaty rocks, and are metamorphic granite or gneiss.

c. Serpentine. We have before mentioned the mineral serpentine. The rock of this name is composed of the minerals steatite, chlorite, iron, hornblende, limestone, &c.

Then come the rocks in which hornblende preponderates. We will only put down hornblende-slate. This is mainly composed of hornblende and felspar. It is a stratified rock and is metamorphic granite.

We pass to the augitic rocks, or those in which augite preponderates.

a. Greenstone. This is *mainly* composed of augite and felspar, the crystals of which are large and distinct.

b. Basalt. This is also mainly composed of augite and felspar; but the crystals are so small that they are indistinct, and the rock presents a homogeneous appearance.

Both these rocks are unstratified; and, instead of being metamorphic granite, have been ejected by volcanoes, and belong to the class of igneous rocks.

Argillaceous rocks are those in which clay or alumina predominates, and are metamorphic from older rocks rich in felspar; but they contain many other minerals besides alumina ones. They are always stratified rocks, and are formed by the disintegration of the particles of older ones.

Calcareous rocks are characterised by containing a large portion of carbonate of lime, which is derived from the skeletons or shells of the inhabitants that have dwelt among them while they were forming. The rest of their constituents are metamorphic; *i. e.* derived from the destruction of older rocks. The most important of these rocks are—

a. Primitive Limestone. This is mainly characterised, in a physical point of view, by its toughness.

b. Secondary Limestone. This is mainly characterised, in a physical point of view, by its easy frangibility.

c. Oolitic Limestone. This is remarkable by its being composed of grains agglutinated together. These grains vary in size from an extreme minuteness to that of a pea.

The minerals chalk, dolomite, and gypsum, before mentioned, are also rocks.

Then there is a class of rock of a still more compound nature than common rocks; to wit, rocks evidently composed of fragments of two or more other rocks. Many of the so-called sandstones are clearly of this nature: and another very mixed rock is the graywacke, which will be noticed in the next